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Chemical Equilibrium Focusing on Acid-Base Systems

Unit D



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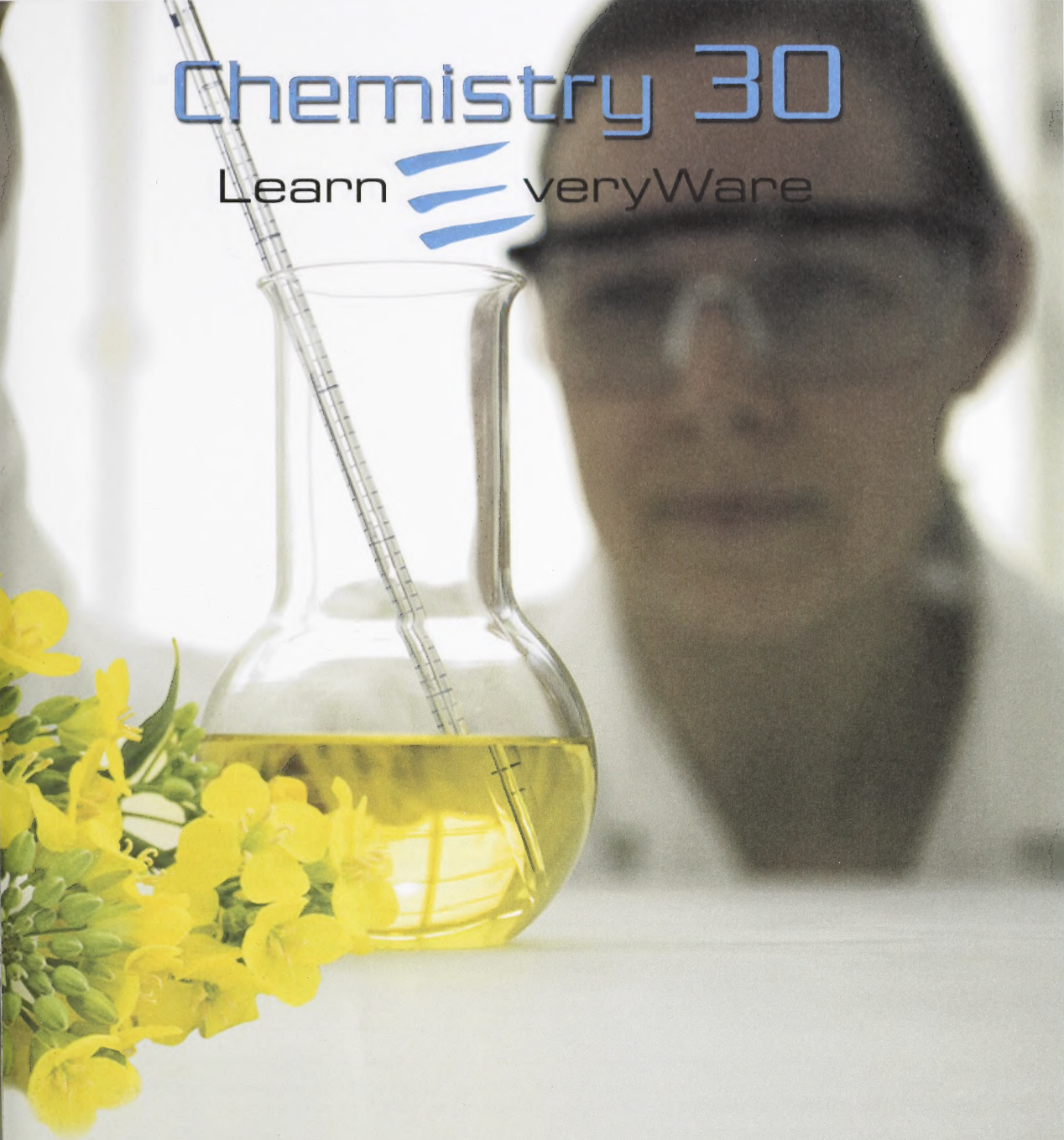
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Chemical Equilibrium Focusing on Acid-Base Systems

Unit D

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Unit D: Chemical Equilibrium Focusing on Acid-Base Systems
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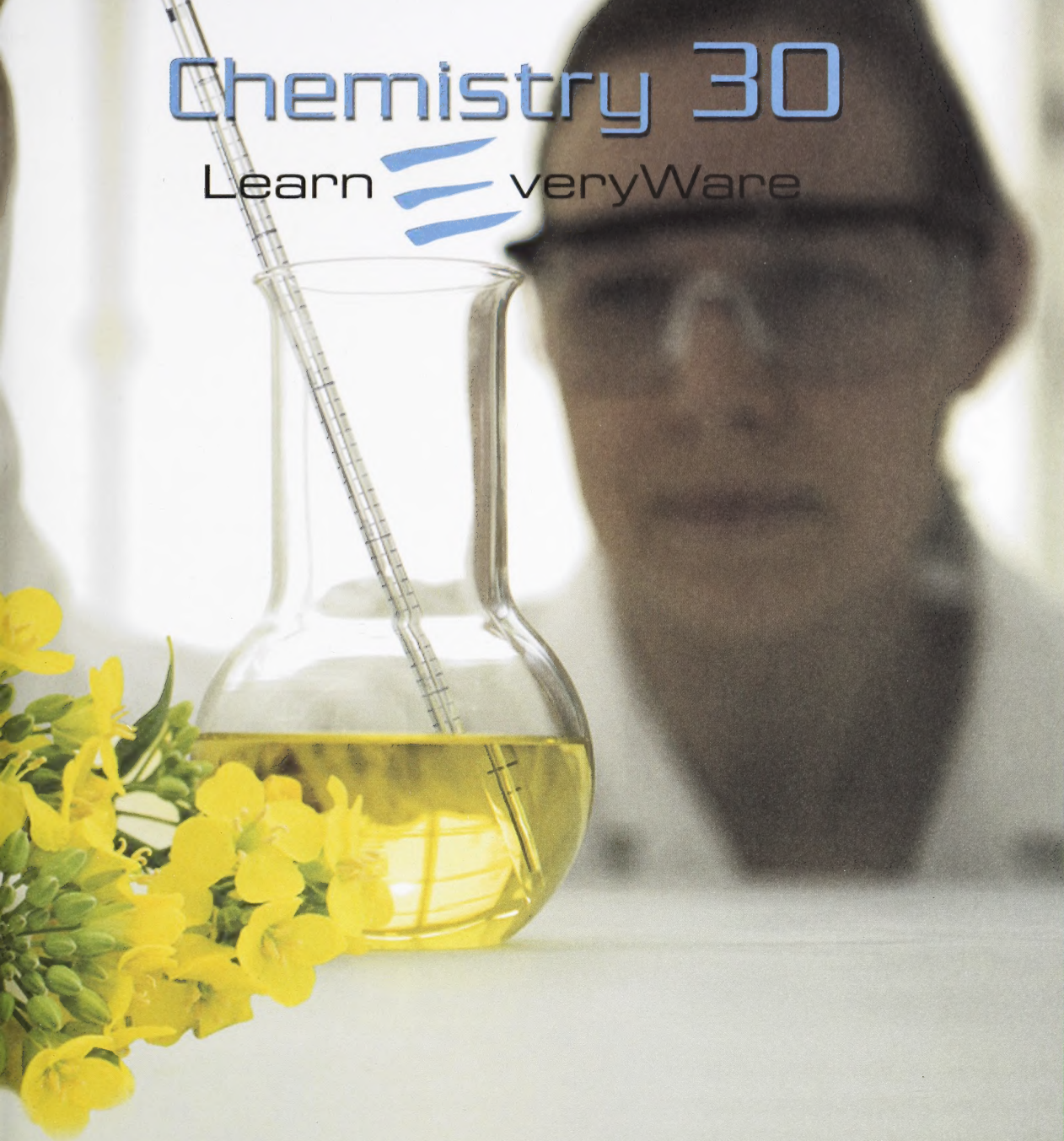
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Principles of Chemical Equilibrium

Module 7

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Unit D Introduction

Unit D is dedicated to a study of chemical equilibrium. In Module 7 you will investigate the principles of chemical equilibrium. In Module 8 you will expand upon these principles as you investigate the equilibrium of acid-base systems.

Throughout Unit D you will rely on your skills as an observer, and you will develop skills to predict changes that may occur to

equilibrium systems. As you have done in other units in this course, you will also rely on your knowledge of diagnostic tests to check your predictions and to confirm your understanding of principles affecting chemical equilibrium.

You will learn to describe and analyze equilibrium systems from a range of perspectives. In addition to qualitative descriptions, you will describe equilibrium systems from a quantitative perspective.

In Module 7 you will use ICE tables and the equilibrium constant, K_c , to describe an equilibrium from a quantitative perspective. You will further apply this understanding in Module 8 using the equilibrium constants for acids and bases, K_a and K_b .

At the end of Unit D you will be able to

- explain that there is a balance of opposing reactions in chemical equilibrium systems
- determine quantitative relationships in simple equilibrium systems

The Module 8 Assessment will also serve as your Unit D Assessment. In the Assessment, you will investigate the principles involved in the design and structure of the Table of Relative Strengths of Acids and Bases, which you will use extensively throughout Module 8.



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Concept Organizer

Unit D—Chemical Equilibrium Focusing on Acid-Base Systems

Unit D—Chemical Equilibrium Focusing on Acid-Base Systems

Module 7—Principles of Chemical Equilibrium

Lesson 1—Introduction to Equilibrium
 Lesson 2—Dynamic Equilibrium
 Lesson 3—Position of an Equilibrium
 Lesson 4—Quantitative Perspectives of Equilibrium
 Lesson 5—The Equilibrium Constant
 Lesson 6—Le Chatelier's Principle

Module 8—Acid-Base Equilibrium

Lesson 1—Introduction to the Equilibrium of Acids and Bases
 Lesson 2—Acid Strength and Equilibrium
 Lesson 3—Bronsted-Lowry Theory
 Lesson 4—Predicting AB Equilibrium
 Lesson 5—Equilibrium Law and the Strength of Acids and Bases
 Lesson 6—pH Curves
 Lesson 7—Buffers

Module Descriptions

Module 7—Principles of Chemical Equilibrium

In Module 7 you will investigate the principles of chemical equilibrium. As you proceed through this module, you will investigate many chemical systems. You will come to understand many characteristics of equilibrium, and you will learn how systems in equilibrium behave. You will be asked to pay special attention to specific properties of each of the systems you investigate.

You will consider the following module questions:

- What is happening in a system at equilibrium?
- How do scientists predict shifts in the equilibrium of a system?

Module 8—Acid-Base Equilibrium

In Module 8 you will apply your understanding of the principles of equilibrium to understanding chemical systems containing aqueous acids and bases. You will use your knowledge of dynamic equilibrium, equilibrium position, and shifts in equilibrium. You will also rely on your ability to analyze an equilibrium system using quantitative techniques (like calculating a value for the equilibrium constant, K_c) and your ability to complete an ICE table.

You will consider the following module question:

- How does the equilibrium of acids and bases affect environmental and biological systems?

Using the Concept Organizer

Each module and lesson contains focusing questions intended to guide your study. Concept organizers are provided for each module. The Module 7 and Module 8 concept organizers, which list the module and lesson questions, are provided at the end of this section.

As you work your way through each module, think about how the lesson questions relate to the module questions and to questions from other lessons. Use the concept organizers to build a concept map or a graphic organizer for each module. To do this, you might use software you already have on your computer, or you might do an Internet search for free software you can use. In your concept map or graphic organizer, record and demonstrate how the lesson and module concepts are connected and interdependent.

Sample concept maps are provided in the Module Summary for each module. Remember that the samples are just that—they show only one of many possible descriptions. However, if your completed concept map or graphic organizer differs significantly from the sample, you may wish to contact your teacher or to compare your map or organizer to those of your classmates. This will ensure that your interpretations of lesson materials and your descriptions are accurate.

Module 7 Concept Organizer

L1 What is equilibrium?

L1 What is happening in a system that is in equilibrium?

L6 What is Le Le Châtelier's principle and how is it used to predict changes in chemical equilibrium systems?

L1 How can you determine whether a system is open or closed?

Module 7 Questions

What is happening in a system at equilibrium?

L2 What is a reverse reaction and why is it important in an equilibrium system?

How do scientists predict shifts in the equilibrium of a system?

L5 What information about an equilibrium is provided in an equilibrium constant?

L2 What changes occur during dynamic equilibrium?

L5 What is the equilibrium constant?

L3 Are all equilibrium states similar?

L4 What procedures are used to calculate the chemical quantities for all substances in a chemical system?

Module 8 Concept Organizer

L1 What is the hydronium ion and how does it account for the properties of acidic solutions?

L7 What is buffering capacity?

L7 What is a buffer and how does a buffer exhibit the characteristics of an equilibrium system?

L6 What events occur in the reaction of polyprotic acids and bases?

L1 What chemical components are part of the equilibrium of water?

L6 What information about acids and bases and their equilibrium is contained on a titration curve?

L1 How does the equilibrium of water affect the calculation of pH?

Module 8 Questions
How does the equilibrium of acids and bases affect environmental and biological systems?

L5 How are values for K_a and K_b used to calculate the pH of solutions containing weak acids and bases?

L2 How is knowledge of equilibrium used to explain the differences in acid strength?

L5 How do K_a and K_b explain the position of the equilibrium of aqueous acids and bases?

L3 What occurs during a chemical reaction between an acid and a base?

L3 How does the Brønsted-Lowry theory support what is known about the equilibrium of aqueous acids and bases?

L3 How do conjugate acid-base pairs represent an equilibrium system?

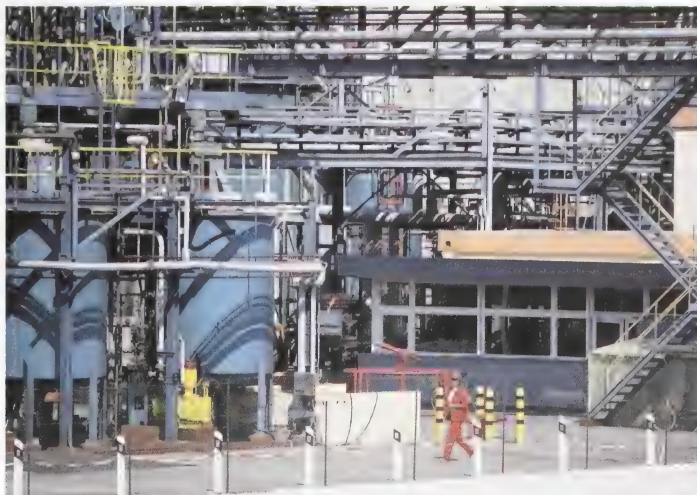
L5 What are K_a , K_b , and K_w ?

L4 How is knowledge of the equilibrium of aqueous systems involved in predicting the outcome of acid-base reactions?

Module 7—Principles of Chemical Equilibrium

Module Introduction

In Module 7 you will examine chemical equilibrium. Chemical equilibrium is something you may not have thought much about, but it is present and it affects many of the chemical systems you have investigated in this course. For example, in the last module you prepared esters. When you cooled the esters on ice, you were manipulating an equilibrium so that you would be better able to collect and observe the ester you had synthesized.



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Other examples of chemical equilibrium are found in biological systems, chemical interactions, and industrial applications. In this module you will investigate many systems and how they demonstrate the principles of systems at equilibrium. You will learn how equilibrium can be classified, studied, and manipulated. You will learn how to predict changes to equilibrium when a system that appears to remain constant is stressed, and you will learn the scientific principles involved when systems are stressed.

In Module 7 you will investigate the following questions:

- What is happening in a system at equilibrium?
- How do scientists predict shifts in the equilibrium of a system?

As you approach the end of the module, you will come to understand many aspects of equilibrium and how systems in equilibrium behave. You will be asked to pay special attention to specific properties of each of the systems you investigate.

Remember that each lesson will also be organized around questions intended to guide your study. As you proceed through Module 7, you may record answers to these questions and any interrelationships that exist between them in a concept map or graphic organizer. More information is available in the Unit D Concept Organizer. In the Module 7 Summary you will receive further information on how you can use your concept map or graphic organizer to review the concepts you studied in this module.



Big Picture

In the Module Introduction you read that there are examples of equilibrium all around you. You are probably questioning whether this is true. In this module you will be introduced to a number of chemical systems in which equilibrium is critical to each system's function.



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Studying equilibrium is not only a matter of knowing the system; studying equilibrium also involves knowing the changes that can occur within the system that help to describe the equilibrium. In your study you will learn about forward and reverse reactions, conditions, and stresses; all of which define a chemical equilibrium. You will do more than study the basic principles of a system at equilibrium; you will learn how manipulating conditions can allow you to understand a chemical equilibrium.

Keep detailed notes. You will use your notes to complete the Module Assessment. In Reflect on the Big Picture activities at the end of each lesson, you will be prompted to construct and add to a table that lists the different systems you investigate. As you update your table you will revisit the concepts you learned in previous lessons, and you will apply that knowledge to new systems. This is a great way to keep using the knowledge you gain in each lesson. You will also be able to use this table to prepare for assessments.



Assessment in This Module

Each lesson contains a range of activities and assessment options. These include assignments, labs, and Self-Check, Try This, Discuss, Reflect and Connect, and Reflect on the Big Picture activities. Instructions will be provided for each of these activities so that you can appropriately focus your time and effort. Your teacher will tell you which assessment options to complete, and which responses to submit for marks or feedback. Remember to save all of your work in your Chemistry 30 folder.

In the Module 7 Assessment you will analyze an equilibrium system, and you will design an experiment to further investigate the factors that influence its equilibrium.

More information is provided in the Module Assessment. You may wish to look at the Module Assessment and the Unit Assessment before starting Lesson 1.

In This Module

Lesson 1—Introduction to Equilibrium

In Lesson 1 you will learn about the two types of equilibrium and how to differentiate between them. You will learn to identify the characteristics of chemical systems that are at equilibrium.

You will investigate the following lesson questions:

- What is equilibrium?
- What is happening in a system that is in equilibrium?
- How can you determine whether a system is open or closed?

Lesson 2—Dynamic Equilibrium

In Lesson 2 you will learn about the three types of equilibrium for chemical systems, and you will learn to identify forward and reverse processes in each.

You will investigate the following lesson questions:

- What is a reverse reaction and why is it important in an equilibrium system?
- What changes occur during dynamic equilibrium?

Lesson 3—Position of an Equilibrium

In Lesson 3 you will learn how an equilibrium can be established and what the relative proportions of products to reactants can be in a system at equilibrium.

You will investigate the following lesson question:

- Are all equilibrium states similar?

Lesson 4—Quantitative Perspectives of Equilibrium

In Lesson 4 you will learn to apply your skills in performing stoichiometric calculations to analyze chemical systems in equilibrium.

You will investigate the following lesson question:

- What procedures are used to calculate the chemical quantities for all substances in a chemical system?

Lesson 5—The Equilibrium Constant

In Lesson 5 you will learn about the equilibrium constant and the equilibrium law. The equilibrium law is an expression that relates the concentrations of the products to the reactants in a chemical system. It is used to calculate an equilibrium constant. You will come to understand how the value for the equilibrium constant is used to communicate information about the position of an equilibrium.

You will investigate the following lesson questions:

- What is the equilibrium constant?
- What information about an equilibrium is provided in an equilibrium constant?

Lesson 6—Le Châtelier’s Principle

In Lesson 6 you will learn how chemical systems at equilibrium react to the application of a stress. The stress can be a change to the concentration of one of the species in the system, a change in the system’s temperature, or a change in the volume (pressure) of the system. You will learn how to predict shifts in equilibrium and to identify the purpose of predicting and manipulating equilibrium by causing shifts.

You will investigate the following lesson question:

- What is Le Châtelier’s principle and how is it used to predict changes in chemical equilibrium systems?

Module 7—Principles of Chemical Equilibrium

Lesson 1—Introduction to Equilibrium



Get Focused

It's one of the last long weekends of the summer and you are spending the weekend beside a lake. Unfortunately, a rain shower is approaching, and it looks like you will have to take a break from your lakeside activities. Fortunately, there is plenty to do inside the cabin with your friends. Maybe you will get a seat at the card game that started a few hours ago.



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It might not seem immediately obvious, but systems at equilibrium are all around you. Can you identify any systems at equilibrium in the situation described in the paragraph above?

To identify a system at equilibrium, you have to know the characteristics that define equilibrium. In this lesson you will begin to investigate equilibrium in chemical systems by learning about the different types of equilibrium that can exist. You will also learn the conditions that define a system at equilibrium.

Consider the following questions as you complete Lesson 1:

- What is equilibrium?
- What is happening in a system that is in equilibrium?
- How can you determine whether a system is open or closed?



Module 7: Lesson 1 Assignment

There is no assignment for this lesson. However, you will begin a table to track the many chemical systems you will investigate and what you learn about their equilibrium. You will add to your table throughout Module 7, and you will use your table to complete the Module Assessment. You will be prompted to begin your table in the Reflect on the Big Picture activity towards the end of this lesson.

The other questions in this lesson are not marked by the teacher; however, you should still answer these questions. The Self-Check, Try This, and other types of questions are placed in this lesson to help you review important information and build key concepts that may be applied in future lessons. You should record the answers to all the questions in the lesson and place those answers in your course folder.

After a discussion with your teacher, you must decide what to do with the questions that are not part of your assignment. For example, you may decide to submit the responses to Try This and other questions that are not marked to your teacher for informal assessment and feedback. Your answers are very important to your teacher. They provide your teacher with information about your learning, and they help your teacher identify where adjustments to your instruction may be necessary.



Explore

What defines an **equilibrium**? You may be surprised to learn that an equilibrium is defined as a system that has constant properties and appears not to change. In the lakeside example from Get Focused, the water level of the lake and the quantity of oxygen dissolved in the lake's water are two examples of equilibria that exist. Can you think of other examples of systems around you that appear not to change and may be equilibrium systems?

Reconsider the example of the lakeside cabin. Either during a rainstorm or at night when all your friends have come inside and no one leaves or enters, an equilibrium has established in the cabin. The equilibrium has established because the number of people in the cabin remains constant. Because no one can come or go, the system (the cabin and its occupants) is a **closed system**.



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Equilibrium can appear in one of two ways: as a **static equilibrium** or as a **dynamic equilibrium**.

During the evening, the position of people inside the cabin could resemble a dynamic equilibrium. This would be true if the number of people playing cards inside the cabin, at any given time, were to remain the same. Over the evening, people could enter and leave the card game, but the number of people at the table at any given time would have to remain constant. As you can see from this example, in a dynamic equilibrium there is a change, but the change is balanced by two processes. In this example the balance is with the number of people entering

and leaving the card game: for every person that enters, one person must leave so that the number of people in the game remains constant.

equilibrium: the state of a closed system in which the system does not appear to change

For a chemical system, physical properties, like colour, remain constant and concentrations of substances within the system do not change.

static equilibrium: the idea that forces on an object are opposing and equal

The object remains constant because no changes are happening internally.

dynamic equilibrium: the idea that there is a balance between two opposing processes (forward and reverse) occurring at the same rate

The balance between the forward and reverse processes maintains the system's appearance of constant properties, in spite of change that has occurred internally.

closed system: a system that is separated from its surroundings by a definite boundary so that energy can enter and leave the system but matter cannot

macroscopic: can be observed with the unaided eye or other senses

microscopic: cannot be observed with the unaided eye or other senses

Now consider the lakeside cottage system more broadly. Consider the closed system to include the cabin, its immediate surroundings, and the number of people not only in the cabin, but in the immediate surroundings as well. In this broader system, no one comes or goes. As such, it is a static system. In the example of this broader system, it would be possible to have a change in **microscopic** properties without having a change in **macroscopic** properties.

You may also have noticed that these examples describe not only static and dynamic equilibrium, but they also describe two different equilibrium states. For example, the number of people expected to be in the cabin when the weather is sunny, is much different than it would be during a rainstorm or at night. Each of these situations is an equilibrium because of the observation of a constant number of people inside the cabin. However, these two equilibrium states are different. Can you think of a reason to account for the difference between the two equilibrium states?

All systems, including those at equilibrium, are influenced by external forces and conditions. Both rain and night are conditions that appear to encourage people to congregate inside the cabin. Under these conditions you can expect one equilibrium state. When the conditions change, so can the equilibrium.

Can you predict what might happen to the equilibrium situation of people in the cabin if you expect more friends to

open system: a system in which both matter and energy can enter and leave the system

arrive throughout the day? As you might expect, adding more people means that your system is no longer closed. An **open system** is one in which matter can be exchanged with its surroundings. In an open system it is much more difficult to predict changes in equilibrium, since it is more difficult to isolate the matter you wish to study.



Read

Read pages 674–676 in the textbook to learn more about the different types of equilibrium and systems.



Try This

Do All Reactions Go to Completion?

Purpose

In this investigation you will observe videos of two chemical systems, and you will write reaction equations to describe the chemical changes you observe.



Procedure

Step 1: Go to the course multimedia DVD, and view “Reaction 1: $\text{HCl(aq)} + \text{Mg(s)}$ ” to observe the components of the system and the change that occurs. Write a balanced chemical equation to describe the changes observed. You may wish to view the video more than once to obtain all the information you need.

Step 2: Review the exploration “Shakin’ the Blues” on page 675 of the textbook. In this system the progress of the chemical reaction is monitored using the colour of the system. Methylene blue indicator in its oxidized form (which can be represented as MBox) appears blue. In its reduced form (which can be represented as MBred), methylene blue indicator appears colourless.

Step 3: Go to the course multimedia DVD, and view “Reaction 2: Shaking the Blues” to observe the components of the system and the change that occurs. Write a balanced chemical equation to describe the changes observed. As you watch the video, observe changes to the colour of the system. What do your observations suggest about the completeness of the reaction and your understanding of chemical equilibrium?

Analysis

TR 1. Which reaction(s) proceeded to completion?

TR 2. How did you describe whether a reaction proceeded to completion in the chemical equations you wrote? How did you describe a reaction that did not appear to be complete?

TR 3. Which system(s) demonstrated that an equilibrium had been established? Explain your reasoning.

Save your responses in your course folder, and submit a copy to your teacher.



Reflect on the Big Picture

In this lesson you have examined some chemical systems, and you have learned about the criteria that define an equilibrium system. In the remaining Module 7 lessons you will investigate more chemical systems and other aspects of equilibrium.

Create a table to summarize what you have learned about the systems so far. You will add to this table throughout Module 7, and you will use your table for the Module Assessment.

Procedure

RBP 1. Create a table like the following:

System	Open or Closed?	Observation of Macroscopic Changes?	Observation of Microscopic Changes?
Cabin and Friends			
HCl(aq) + Mg(s)			
Methylene Blue, Glucose, Potassium Hydroxide, and Oxygen			

RBP 2. Fill in the table with the systems you investigated in this lesson and your observations for each system.



Reflect and Connect

RC 1. Take some time as you go through your day to note systems in your life that are examples of either static or dynamic equilibrium. Choose the best examples and describe why each one meets the criteria for a system at equilibrium. Summarize your examples on your table of equilibrium systems.



Post your examples of systems at equilibrium to the discussion area for your class. Review the postings of at least two other students. Provide feedback to help the students improve the precision of their examples. Revise your examples using feedback you receive from other students. Save your revised examples in your course folder.



Lesson Summary

In Lesson 1 you considered the following questions:

- What is equilibrium?
- What is happening in a system that is in equilibrium?
- How can you determine whether a system is open or closed?

You learned that equilibrium is defined as a system having constant properties and appearing not to change. There are two main types of equilibrium: static equilibrium and dynamic equilibrium. In a dynamic equilibrium there is a balance between two opposing processes, the forward and reverse reactions, both occurring at the same rate. You also learned to differentiate between a closed system and an open system.

You applied your newly acquired knowledge to identify examples of equilibrium systems. As you work through Module 7 you will continue to examine these systems to describe their other attributes, and you will consider how these attributes are similar in all equilibrium systems.

Lesson Glossary

closed system: a system that is separated from its surroundings by a definite boundary so that energy can enter and leave the system but matter cannot

dynamic equilibrium: the idea that there is a balance between two opposing processes (forward and reverse) occurring at the same rate

The balance between the forward and reverse processes maintains the system's appearance of constant properties, in spite of change that has occurred internally.

equilibrium: the state of a closed system in which the system does not appear to change

For a chemical system, physical properties, like colour, remain constant and concentrations of substances within the system do not change.

macroscopic: can be observed with the unaided eye or other senses

microscopic: cannot be observed with the unaided eye or other senses

open system: a system in which both matter and energy can enter and leave the system

static equilibrium: the idea that forces on an object are opposing and equal

The object remains constant because no changes are happening internally.

Module 7—Principles of Chemical Equilibrium

Lesson 2—Dynamic Equilibrium



Get Focused

In Lesson 1 you observed the colour changes of methylene blue that resulted when the system the indicator was in was shaken or allowed to sit undisturbed. In your analysis of that system you learned that chemical reactions are reversible.

To some extent all chemical systems demonstrate reversibility. Therefore, studying the equilibrium of chemical systems involves recognizing the forward and reverse reactions for a system and their reaction rates. For example, if the evaporation of water into the atmosphere is considered the forward reaction for the water in the biosphere system, then precipitation as rain, snow, or fog would be the products of the reverse reaction.

You may recall that you learned about reaction rate in Unit A when you learned about the collision-activation theory, which explains the events involved in a chemical reaction. How might the probability of collisions between particles and energy of activation influence a chemical system at equilibrium?



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In this lesson you will look for evidence of the forward and reverse reactions that occur within chemical systems. By accumulating and interpreting this evidence, you will gain a better understanding of dynamic equilibrium.

Consider the following questions as you complete Lesson 2:

- What is a reverse reaction and why is it important in an equilibrium system?
- What changes occur during dynamic equilibrium?



Module 7: Lesson 2 Assignment

You will be directed to go to the Assignment Booklet to complete the Module 7: Lesson 2 Assignment later in this lesson. You will also receive further information on how to add to the table you began in Lesson 1.

You must decide what to do with the questions that are not marked by the teacher.

Remember that these questions provide you with the practice and feedback that you need to successfully complete this course. You should respond to all the questions and place those answers in your course folder.



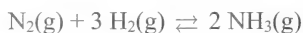
Explore



Try This

The Haber process to make ammonia provides an important example of the importance of forward and reverse reactions in a chemical system. Read “The Haber Process” on page 325 in the textbook.

The problems encountered when trying to increase the yield of ammonia are evidence that a reverse reaction is an important aspect of this system. It can occur at the same instant as the forward reaction and in the same reaction vessel.



In designing a system to maximize the yield of ammonia, experimentation with pressures, temperatures, and catalysts is intended to favour the forward reaction so that large quantities of product can be made. During unsuccessful trials, large quantities were not achieved. In Lesson 6 you will learn more about conditions that may favour the forward or reverse reaction. For now, it is important to understand that both processes exist within a system.

In Lesson 1 you learned that a system at equilibrium is a closed system that has constant properties or does not appear to change. In this lesson you have learned that an equilibrium also involves a forward and a reverse reaction. How do all of these ideas connect? Use the following activity to demonstrate your understanding of these aspects of an equilibrium.



Try This

Use the following statement to answer questions TR 1–3: A system at equilibrium is a closed system that has constant properties.

TR 1. Interpret the statement by describing the rate of reaction of the forward and reverse processes for a system that is at equilibrium.

TR 2. Interpret the statement by describing the rates of the forward and reverse reaction for a system that is approaching equilibrium.

TR 3. Provide one observation to support your answer to TR 2.

Save your work in your course folder and submit a copy to your teacher.

**Read**

Read page 677 in the textbook to learn more about the different types of equilibria that can exist in chemical systems.

**Self-Check**

SC 1. Provide an example of each type of chemical equilibrium described on page 677 of the textbook. Identify the chemical components involved.

Check your work with the answer in the appendix.

**Module 7: Lesson 2 Lab—Evidence of Reverse Reaction****Purpose**

In your chemistry study to date, you have studied reactions that are quantitative. What does a quantitative reaction mean? How does the term *quantitative* connect with what you are learning about equilibrium? In this virtual investigation, you will observe the results of tests performed on a chemical system to learn more about equilibrium.

As you read the experimental design, you may wish to write out the chemical reactions involved and determine what substance is being tested for using $\text{Ba}(\text{NO}_3)_2(\text{aq})$. Contact your teacher if you are having difficulty.

Problem

Can the products of the reverse reaction be detected in a chemical system at equilibrium?

Experimental Design

Four different volumes (25 mL, 50 mL, 75 mL, and 100 mL) of 0.50 mol/L $\text{CaCl}_2(\text{aq})$ are added to four 50-mL aliquots of 0.50 mol/L $\text{Na}_2\text{SO}_4(\text{aq})$. Each system is filtered and the mass of precipitate produced is dried and measured. The filtrate from each system is tested with a saturated solution of $\text{Ba}(\text{NO}_3)_2(\text{aq})$.



Open the Assignment Booklet, and turn to "Module 7: Lesson 2 Assignment." Answer the Pre-Lab and Hypothesis questions. Record data in the Assignment Booklet as you view the following lab.



Go to the course multimedia DVD, and view "Lab—Evidence of Reverse Reaction."

Analysis



Module 7: Lesson 2 Assignment

You will complete the Analysis in your Assignment Booklet. Open the Assignment Booklet and complete “Module 7: Lesson 2 Assignment” now.



Reflect on the Big Picture

In this lesson you examined a number of chemical systems, and you learned about the forward and reverse reactions that describe each system’s dynamic equilibrium. In Lesson 1 you prepared a table to keep a record of the equilibrium systems you investigate throughout this module.

RBP 1. Add the systems that you looked at in this lesson to your table.

RBP 2. Add two columns to your table. Use the headings “Static or Dynamic Equilibrium?” and “Type of Chemical Equilibrium.” (You may wish to adjust the set-up for the page so that it is landscape, since you will be adding additional columns in future lessons.)

RBP 3. Review all of the systems listed in your table, and complete all empty cells in your table.

Save your updated table in your course folder.



Lesson Summary

In Lesson 2 you considered the following questions:

- What is a reverse reaction and why is it important in an equilibrium system?
- What changes occur during dynamic equilibrium?

You learned that there are three types of equilibrium for chemical systems:

- Phase equilibrium: one chemical substance exists in more than one phase in a closed system.
- Solubility equilibrium: one solute and one solvent exist in a saturated solution that has an excess of solute in a closed system.
- Chemical reaction equilibrium: the reactants create products as fast as the products create reactants; in other words, the forward reaction is approximately equal to the reverse reaction.

Each of these types of equilibria is a dynamic equilibrium, meaning that microscopic changes occur in the system as described by forward and reverse reactions for the chemical components of the system. In this lesson you completed an investigation to test for the products of the reverse reaction. This investigation confirmed that this process occurs, even in conditions involving stoichiometric excess.

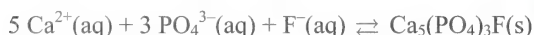
Module 7—Principles of Chemical Equilibrium

Lesson 3—Position of an Equilibrium



Get Focused

When you brush your teeth you probably do so to remove harmful plaque that can cause tooth decay. Did you realize that brushing your teeth contributes to an equilibrium system? Most types of toothpaste and dental rinse contain fluoride, an important component in the process of building dental enamel—the outermost tooth layer made of hydroxyapatite and fibrous protein. The layer of enamel is part of a chemical equilibrium involving the processes of enamel formation and erosion. The chemical equilibrium is shown in the reaction



Can you explain why dentists often ask their patients to use toothpastes and dental rinses containing fluoride? How would you incorporate your understanding of equilibrium to explain this practice?

In Lesson 2 you learned about the three types of chemical equilibria. In Lesson 3 you will study chemical systems in equilibrium in greater depth, and you will investigate the characteristics of these systems.

Consider the following question as you complete Lesson 3:

- Are all equilibrium states similar?



Module 7: Lesson 3 Assignment

In your lesson assignment you will analyze a model of dynamic equilibrium. You will be directed to go to the Assignment Booklet to complete the Module 7: Lesson 3 Assignment later in this lesson. You will also receive further information on how to add to the table you started in Lesson 1.

You must decide what to do with the questions that are not marked by the teacher.

Remember that these questions provide you with the practice and feedback that you need to successfully complete this course. You should respond to all the questions and place those answers in your course folder.



Explore



Read

Read the mini investigation “Modelling Dynamic Equilibrium” on page 678 of the textbook.



Go to the course multimedia DVD, and view “Mini Investigation: Modelling Dynamic Equilibrium.” As you view the presentation, record the volume of liquid in each cylinder after each transfer of liquid using the straws. Obtain the data showing the results of three other trials for this experiment.



Module 7: Lesson 3 Assignment

Open the Assignment Booklet. Complete “Module 7: Lesson 3 Assignment.”



Read

Did all four systems you examined in “Mini Investigation: Modelling Dynamic Equilibrium” achieve an equilibrium?

As you learned earlier, equilibrium in a chemical system exists when the rates of the forward and reverse reactions equal. As you saw in this

investigation, the initial

volume of reactant, relative to product, was not important in establishing an equilibrium.



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Another important aspect of equilibrium is position of equilibrium. The term *equal* brings to mind a teeter-totter balancing parallel to the ground. If each side of the teeter-totter represents the concentrations of reactant and products respectively, is a position parallel to the ground the only position possible? If you think about some of the other equilibrium systems you have investigated and summarized in the table you have been keeping, does their equilibrium represent equal concentrations of reactant and products?

Principles of Chemical Equilibrium

In the “Evidence of a Reverse Reaction” lab you completed in Lesson 2, you will recall that the equilibrium in that system involved a great deal of calcium sulfate precipitate and an extremely small quantity of dissociated sulfate ions, which were only detectable by the addition of barium ions. This equilibrium clearly favoured the formation of the products (forward reaction) and resulted in a situation in which the concentration of the products was higher than that of the reactant species.

Systems that significantly favour the formation of products are quantitative systems. You will recall from your work in Chemistry 20 that you were able to perform stoichiometric calculations on chemical systems that had percentage reactions of over 99%. In these systems and relative to the quantity of product, very little reactant remains once equilibrium is established.

Read the section “Chemical Reaction Equilibrium” on pages 678–680 in the textbook to learn more about the different positions of equilibrium in chemical systems.



Self-Check

SC 1. Complete “Practice” questions 3, 4, and 7 on page 682 in the textbook.

Check your work with the answer in the appendix.



Reflect and Connect

The addition of fluoride to tooth care routines has significantly reduced the amount of tooth decay. You will find fluoride in two main sources: in your toothpaste and in your drinking water. In most cities in Alberta, fluoride is added to the drinking water in tiny doses to help fight tooth decay. In some well water, fluoride is found in much greater concentrations than in treated city water. If a little fluoride is helpful to build tooth enamel, then is more better? If you look at the information on a tube of toothpaste, you will see a warning not to ingest your toothpaste. Although a little fluoride can have excellent benefits, too much fluoride can have negative health effects.



Reflect on the Big Picture

In this lesson you have examined a number of chemical systems, and you have learned about the forward and reverse reactions that describe their dynamic equilibrium. In Lesson 2 you added to your table of equilibrium systems investigated in this module.

RBP 1. Add the systems that you looked at in this lesson to your table.

RBP 2. Add an additional column to your table. Make the heading for the new column “Position of Equilibrium.”

RBP 3. Review all of the systems listed in your table and complete all empty cells.

Save your updated table in your course folder.



Lesson Summary

In Lesson 3 you considered the following question:

- Are all equilibrium states similar?

You completed a simulation of a system demonstrating dynamic equilibrium. This investigation helped you learn more about how an equilibrium can be established and what the relative proportions of products to reactants can be in a system at equilibrium.

You also learned that there are four positions for an equilibrium involving a chemical system:

- Non-spontaneous: these systems have no apparent reaction.
- Products favoured: the equilibrium has a percent yield that is greater than 50%, where the products of the forward reaction are favoured. The outcome for such a system is that the concentration of the products will be higher than that of the reactant species. You need to be able to differentiate this from what occurs when the system reaches its equilibrium and the rates of the forward and reverse reactions in the system are equal.
- Reactants favoured: the equilibrium has a percent yield that is less than 50%, where the reactants of the forward reaction are favoured. The outcome for such a system is that the concentration of the products will be lower than that of the reactant species.
- Quantitative: the equilibrium has a percent yield that is greater than 99%, where the products of the forward reaction are highly favoured.

Module 7—Principles of Chemical Equilibrium**Lesson 4—Quantitative Perspectives of Equilibrium****Get Focused**

In your study of chemical equilibrium so far, you have observed colour changes. In chemical systems in which a coloured substance is involved, colour changes indicate a change in the molar concentration of the coloured substance within the system. Depending on the concentration of one coloured substance relative to other substances, the overall system's colour describes the position of a system's equilibrium. If the overall colour changes, it indicates that the equilibrium is changing.

Using colour is not always the best way to describe the position of an equilibrium. In Lesson 3 you determined the position of a chemical equilibrium system using percent reaction. To complete a calculation of percent yield, you require information about the chemical quantities of one reactant and one of the products in the system. If a chemical system involves many components, how can their chemical quantities be determined? What information about the system do these quantities provide?



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Determining the quantities of all the reactants and products in a system is an important process in understanding a chemical equilibrium. In Lesson 4 you will learn a systematic procedure to better understand chemical equilibrium from a quantitative perspective.

Consider the following question as you complete Lesson 4:

- What procedures are used to calculate the chemical quantities for all substances in a chemical system?

**Module 7: Lesson 4 Assignment**

There is no assignment for this lesson. You will be assessed on the material you learn in this lesson as part of the Lesson 5 Assignment.

You must decide what to do with the questions that are not marked by the teacher.

Remember that these questions provide you with the practice and feedback that you need to successfully complete this course. You should respond to all the questions and place those answers in your course folder.



Explore



Self-Check

In Lesson 3, you completed a series of calculations to determine the percent yield to analyze the following system:



Initial quantity of $\text{CH}_4(\text{g}) = 2.00 \text{ mol}$

Quantity of $\text{CH}_3\text{Cl}(\text{g}) = 1.40 \text{ mol}$

SC 1. Calculate the quantity of $\text{CH}_4(\text{g})$ that reacted and that remains unreacted. Explain how you completed these calculations.

SC 2. Recall that 10.00 mol of chlorine were placed into the system initially. Could the process you explained in SC 1 be used to calculate the quantity of $\text{Cl}_2(\text{g})$ and $\text{HCl}(\text{g})$ in the system at equilibrium?

Check your work with the answer in the appendix.



Read

The process you described in your answers to SC 1 and SC 2 requires you to apply your knowledge of stoichiometry to a new situation. You know the quantity of one of the products and one of the reactants, and you have a balanced chemical equation. It is possible to use this information and the ratios of coefficients between species to determine the quantities of the other substances in the system.

Using changes in chemical quantities of substances in a system as it approaches equilibrium can be summarized using a structure called an ICE table. To read more about an ICE table and to learn how the calculations you performed are used in its completion, read page 681 in the textbook and work through “Sample problem 15.1.” As you work through the sample problem, note how the stoichiometric ratios are applied to the change in concentration for each species in the system.



Self-Check

SC 3. What do the letters in ICE represent?

SC 4. Why is the value for the change in the quantity of each species important in completing an ICE table?

SC 5. What does a negative sign in the change row of an ICE table represent?

Check your work with the answer in the appendix.



Try This



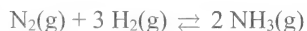
Go to the course multimedia DVD, and open “Virtual ICE Table.” Use the information provided to determine the values for the empty cells in the table. Enter your answers into the appropriate cells to determine whether your answers are correct.



Self-Check

SC 6. Complete “Practice” question 6 on page 682 of the textbook.

SC 7. An experiment to study the equilibrium in the following system is conducted:



In the experiment, 2.42 mol of $\text{NH}_3(\text{g})$ and 20.76 mol of $\text{H}_2(\text{g})$ were added to an evacuated 2.00-L flask. At equilibrium the concentration of ammonia in the flask was 1.05 mol/L.

Prepare a table showing the initial, change, and equilibrium concentrations for each of the substances in the system.

Check your work with the answer in the appendix.



Reflect and Connect



When taking antibiotics, it is advised that you take the drug at regular time intervals. Can you think of why it is important to take the drug at the times specified?

In order for a drug to have its desired effect, it has to reach a certain concentration in the body. This concentration can be regarded as an equilibrium concentration, since the concentration is a result of the processes of ingestion and absorption of the drug (which will increase the concentration) and the processes of metabolism and excretion (which will decrease

the concentration). Can you think of how the processes to reach an equilibrium in the concentration of a medicine in the body are similar to those in the investigation in which you simulated a dynamic equilibrium?



Lesson Summary

In Lesson 4 you considered the following question:

- What procedures are used to calculate the chemical quantities for all substances in a chemical system?

You applied your skills in performing stoichiometric calculations to analyze chemical systems in equilibrium. You found that constructing an ICE table is a convenient way to identify the relevant calculations that must be performed to complete this type of analysis.

In Lesson 5 you will learn more about how an equilibrium system can be analyzed using quantitative techniques.

Module 7—Principles of Chemical Equilibrium

Lesson 5—The Equilibrium Constant



Get Focused

In Lesson 4 you considered the importance of maintaining the desired concentration of a prescription drug in your system. Maintaining the concentration of a drug in a system involves many equilibria, including the rates at which the drug could be absorbed into your system and the rate at which the drug is metabolized and excreted.

A pharmacologist studies the development and action of drugs in the body. Consideration of equilibria is very important in the work of pharmacologists. Important equilibria to consider in pharmacology include the solubility of the drug and its ionization once in the system. These equilibria influence the ability of the drug to be absorbed by the blood and to have its desired effect in the body. How would a pharmacologist and other scientists study an equilibrium? What sort of information is most useful to them?

In this lesson you will learn how to use numerical data, beyond the percentage reaction calculation performed in previous lessons, to provide information about an equilibrium, including the equilibrium's position.



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Consider the following questions as you complete Lesson 5:

- What is the equilibrium constant?
- What information about an equilibrium is provided in an equilibrium constant?



Module 7: Lesson 5 Assignment

Your assignment will address material covered in Lessons 4 and 5. You will be directed to go to the Assignment Booklet to complete the Module 7: Lesson 5 Assignment later in this lesson. You will also receive further information on adding to the table you began in Lesson 1.

You must decide what to do with the questions that are not marked by the teacher.

Remember that these questions provide you with the practice and feedback that you need to successfully complete this course. You should respond to all the questions and place those answers in your course folder.



Explore

In your science and mathematics courses you have come across many “constants.” A constant is a numerical value that describes a relationship between variables within a system. Sometimes constants can have units that help identify the variables they are relating. For example, in previous chemistry courses you were introduced to the universal gas constant, $8.314 \text{ L}\cdot\text{kPa}/(\text{mol}\cdot\text{K})$. You may have even completed an investigation to derive this value from the data you collected.

In the next activity you will derive another familiar constant from your observations—pi.



Try This

Derivation of Pi

Materials

- piece of string 30 cm in length
- ruler
- calculator

Procedure



Step 1: Go to the course multimedia DVD, and then download and print the handout “Try This—Circles.”

Step 2: Prepare a data table consisting of four columns and six rows. Label the columns with the following headings: “Circle Number,” “Circumference, C (cm),” “Diameter, d (cm),” “Circumference Divided by Diameter, C/d .”

Step 3: Using the piece of string and the ruler, measure the circumference and diameter of each circle on the handout. Record each value in your data table.

Step 4: Calculate the circumference divided by the diameter for each circle. Record the value obtained for each circle in the appropriate cell of your data table.

Analysis

TR 1. Comment on the results of the calculations for each circle.

TR 2. Do the results demonstrate that a constant relationship exists between the variables of the circles being analyzed?



Read

The constant π is used in most calculations involving circle geometry. When completing certain calculations involving gases, you may have used the ideal gas constant or the constants for molar volume of a gas at STP or SATP. Do systems at equilibrium each have a constant? What are the variables that describe an equilibrium? How would these variables be used to calculate a constant?

In your earlier investigations of equilibrium, you used colour changes to monitor whether or not a system was at equilibrium. You will recall that in an equilibrium, all species in the chemical system are present to some extent. Therefore, the equilibrium you observed was a mixture of colours of the reactant and product species.

Depending upon the position of the equilibrium, the colour may have been more like the products or the reactants, but it was always a blending of colours from each group. You also know from your understanding of solutions that colour intensity is directly related to concentration. For example, the higher the concentration of the oxidized form of methylene blue, the more intense the blue colour of the solution in the flask.

From what you have learned so far in this module, you know that some relationships regarding equilibrium are important and need to be considered in any calculation for an equilibrium constant.

Keep the following three important relationships in an equilibrium in mind when calculating an equilibrium constant:

- The concentrations of each species in the system are important values.
- The position of an equilibrium is determined by the relative proportions of products to reactants.
- A balanced chemical equation describes the proportions of reactants and products involved in the system.



Try This

An Equilibrium Law

Purpose

To test various mathematical formulas to see if they can provide an equilibrium law, a general formula describing how to calculate an equilibrium constant is required.

TR 3. Read “Lab Exercise 15.A” on page 683 of the textbook. Complete an analysis by calculating an answer for each of the proposed formulas for an equilibrium law. Use the trials shown in “Table 6” on page 683. A suitable formula will produce the same value (a constant) from the data for all five trials.

TR 4. Which of the proposed formulas tested in “Lab Exercise 15.A” provides a “constant” for the data analyzed?

Summarize the results of your calculations in a table. Save a copy of your table in your course folder.



Discuss

Post your answer and the table you developed in TR 3 to the discussion area for your class. Use your classmates’ responses to check your skill in accurately completing the calculations and to see whether you can identify a similar relationship that could be used to calculate an equilibrium constant.

Earlier you were asked to consider the three important relationships that should be incorporated into a mathematical approach to calculating an equilibrium constant. To review, these were as follows:

- The concentrations of each species in the system are important values.
- The position of an equilibrium is determined by the relative proportions of products to reactants.
- A balanced chemical equation describes the proportions of reactants and products involved in the system.

D 1. Comment on whether the proposed formula you identified in TR 4 also incorporates the three important relationships in an equilibrium described above.



Read

Completing the analysis of the data shown in “Lab Exercise 15.A” provides valuable insight into what an equilibrium needs to be. Read from page 684 to the end of the summary on page 686 of the textbook to learn more about the development of an equilibrium law and the equilibrium constant, K_c . Work through the “Communication examples” in this section.



Try This



In this activity, you will work with the simulation “The Meaning of the Equilibrium Constant.” Go to the website for your textbook. You will need your username and password for the textbook. (Remember to ask your teacher if you do not have these.) Once logged in, click on each of the following links in order to navigate to the simulation: Chapter 15 Equilibrium Systems, Section 15.1, Extensions, Simulation, Extension: The Meaning of the Equilibrium Constant. The simulation reinforces the information on page 685 in the textbook.

As you work with this simulation, note that you are able to manipulate the following variables:

- value of K_c , the equilibrium constant
- volume of the container
- temperature of the system
- number of spheres in the system

TR 5. As you view the simulation you will be able to see information that describes the numbers of each particle, the speed of particles in the container, the container size, and the concentration of particles of each type versus time. Which of these pieces of information would provide the greatest information about the equilibrium in the system?

TR 6. Prepare a table to record the changes you wish to observe while manipulating each of the variables listed in the bullet list. Record observations of how the equilibrium is affected by the changes you make.

You may wish to complete the self-scoring quiz that appears in the lower box of the animation window. Make note of answers you do not get correct. You may wish to verify your understanding using the textbook or other reference materials.

Save your work in your course folder.

**Read**

In some situations, the value for K_c can be used to determine the expected change in the concentration of substances, as equilibrium is established. In such a situation it is important to remember that the value you are required to solve for, is the change in concentration. You may wish to use the structure of an ICE table to help define your variables and relationships. This will help you complete the mathematical calculations.

Read “Predicting Final Equilibrium Concentrations” on pages 686–687 of the textbook. Work through “Sample problem 15.2” in the reading to see how these types of problems are solved.

**Self-Check**

SC 1. Complete “Section 15.1” questions 1 (a, b, e, f), 2, 4 (a, b), 5, 6, 7, 9, and 10 on pages 688–689 in the textbook.

Check your work with the answer in the appendix.

**Module 7: Lesson 5 Assignment**

Go to the Assignment Booklet. Complete “Module 7: Lesson 5 Assignment.”

**Reflect on the Big Picture**

In this lesson you have examined a number of chemical systems, and you have learned about the equilibrium law. It's time to add to the table you began in Lesson 1. You last added to your table in Lesson 3.

RBP 1. Add an additional column to your table. Name the column “Equilibrium Expression.”

RBP 2. Write the equilibrium law for each system on the table, or a description of how this would be done.

Save your updated table in your course folder.



Lesson Summary

In Lesson 5 you considered the following questions:

- What is the equilibrium constant?
- What information about an equilibrium is provided in an equilibrium constant?

In this lesson you learned about the equilibrium constant and the equilibrium law. The equilibrium law is an expression that relates the concentrations of the products to the reactants in a chemical system. The equilibrium law is used to calculate an equilibrium constant. The equilibrium law can be stated as

$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b},$$

where the balanced chemical equation for the system takes the form $aA + bB \rightleftharpoons cC + dD$

As you found in this lesson, the equilibrium constant is a specific value for a reaction system that is constant for a full range of concentrations. The magnitude of the value for the equilibrium constant provides information about the relative proportions of the products and reactants in the system:

- if $K = 1$, then $[\text{products}] = [\text{reactants}]$
- if $K > 1$, then $[\text{products}] > [\text{reactants}]$
- if $K < 1$, then $[\text{products}] < [\text{reactants}]$

Module 7—Principles of Chemical Equilibrium

Lesson 6—Le Châtelier's Principle



Get Focused

It's winter and time to get back into the competitive season. The first few races of the snowboarding season are always so tough. Despite training hard in the off-season, snowboarders always feel out of breath and tired when in the mountains.



© Danila/shutterstock

This kind of experience is common for people who visit altitudes, especially if they exert themselves through strenuous physical activity. You may have noticed that people who live at high altitudes do not seem to suffer from these symptoms. The human body's ability to adapt to living at altitude has attracted many athletes to live and train in high-altitude locations or to purchase tents that simulate high-altitude conditions.

Since the quantity of oxygen in the air at higher altitudes is reduced as compared to lower altitudes, the human body has to respond to this situation. If you live under these conditions for a prolonged period of time, your body will produce more red blood cells, the blood component responsible for oxygen transport. By producing more red blood cells, your body can capture more of the oxygen from the air, thereby providing your body with sufficient oxygen.

In Lesson 6 you will learn about responses that occur in chemical systems as a result of changes in the conditions they are within. The change to conditions, often referred to as a stress, can change an equilibrium. In this lesson you will investigate whether there is a way to predict the response to a stress placed on a chemical system in an equilibrium.

Consider the following question as you complete Lesson 6:

- What is Le Châtelier's principle and how is it used to predict changes in chemical equilibrium systems?



Module 7: Lesson 6 Assignment

There is no assignment for this lesson. You will be assessed on the material you learn in this lesson as part of the Module Assessment. Later in this lesson you will be reminded to update the table you began in Lesson 1.

You must decide what to do with the questions that are not marked by the teacher.

Remember that these questions provide you with the practice and feedback that you need to successfully complete this course. You should respond to all the questions and place those answers in your course folder.



Explore



Try This

Demonstration of Equilibrium Shifts



Go to the course multimedia DVD, and view the virtual investigation “Demonstration of Equilibrium Shifts.”

Analysis

TR 1. For each of Tubes 2–5, indicate which substance, $\text{Co}(\text{H}_2\text{O})_6^{2+}$ or CoCl_4^{2-} , increases in concentration as a result of the manipulation performed on the tube during the demonstration.

TR 2. Suggest an explanation for the colour changes you observe.

Save a record of your observations and answers in your course folder. Throughout this lesson you will review your observations and consolidate your understanding of the principles introduced and developed throughout Lesson 6.



Read

In the demonstration you just completed, you saw that changing the conditions for a chemical system equilibrium causes a change to the equilibrium. In the demonstration, a colour change was distinct evidence that a shift in the position of an equilibrium had occurred. In your analysis of the observation, you were asked to identify which coloured species, $\text{Co}(\text{H}_2\text{O})_6^{2+}$ or CoCl_4^{2-} , increased in concentration as a result of the change in conditions on that system.

How can the concentration of one of these substances change? Read pages 691–693 in the textbook to begin to address this question.

Le Châtelier's principle describes a response by a chemical system at equilibrium that has been disturbed. This principle states that when a chemical equilibrium is disturbed, the system seems to respond in the direction that opposes the change (consumes or counteracts the stress). This principle allows you to predict the direction of the shift in the equilibrium.



Discuss

In the virtual investigation “Demonstration of Equilibrium Shifts,” you observed the effect of changing the concentration and temperature on the system being studied. Can you explain the colour change you observed in each of the tubes tested using Le Châtelier's principle?

D 1. Retrieve your observations and answers to TR 1 and TR 2. Review the changes that occurred in each tube, and write an explanation for the change observed that demonstrates an understanding of Le Châtelier's principle. Make sure your explanation for the change that occurred in each tube explicitly identifies the following points:

- the “stress” placed on the system
- which reaction, the forward or the reverse, is favoured to respond to the stress
- how favouring the forward or reverse reaction serves to counteract the stress
- the direction of the equilibrium shift

Once you have completed your answers, post them to the discussion area for your class. Review the postings of at least two other students. Provide feedback to the students to help them make their responses more clear and accurate, and to help them ensure they have addressed the aspects listed above.

Revise your response based on the textbook readings and any feedback you receive, and save your revised response in your course folder.



Read

For equilibrium systems containing gases, a change in the volume of the system can also be a stress that will bring about a shift in equilibrium. Read page 694 of the textbook to learn about the effect that a change in volume and the addition of a catalyst can have on the equilibrium of a system. As you read this section you may wish to recall the pressure–volume relationship you learned about in previous science courses—a halving of the volume of a fixed quantity of gas results in a doubling of the gas' pressure.

Principles of Chemical Equilibrium

There are three main changes that can affect the equilibrium in a chemical system: concentration change, temperature change, and volume (or pressure) change.

- **Concentration change:** An increase in concentration causes a shift to consume some of the added reactant or product. A decrease in concentration causes a shift to produce some of the removed reactant or product.
- **Temperature change:** An increase in temperature of a system causes a shift to absorb some of the added heat energy. A decrease in temperature causes a shift to replace some of the energy that has been removed.
- **Volume (or pressure) change:** An increase in volume (decrease in pressure) causes a shift toward the side with the larger total chemical amount of gaseous entities. A decrease in volume (increase in pressure) causes a shift toward the side with the smaller total chemical amount of gaseous entities.



Self-Check

SC 1. Complete “Practice” questions 2 and 3 on page 695 of the textbook.

Check your work with the answer in the appendix.



Reflect on the Big Picture

In this lesson you have examined a number of chemical systems and the equilibrium of each one. As you have investigated the systems, you have learned how factors such as pressure, change in concentration, and temperature can act as a stress on these systems. Retrieve the table you began in Lesson 1.

RBP 1. Add two additional columns to your table. Make the headings for the new columns “Equilibrium System Stressed By” and “How Stress Shifts This Equilibrium.”

RBP 2. Complete the empty cells in your table for all of the systems listed.

Save your updated table in your course folder.



Lesson Summary

In Lesson 6 you considered the following question:

- What is Le Châtelier's principle and how is it used to predict changes in chemical equilibrium systems?

Le Châtelier's principle describes how chemical systems at equilibrium react to stresses such as changes in concentration of species in the system, change to temperature, and changes in the volume (pressure) of the system. Equilibrium systems undergo a shift to a new equilibrium by either favouring the forward or the reverse reaction. In this lesson you learned how to identify a stress placed on a system, the component in the system affected by the stress, and the most likely action by the system to counteract the stress.

You are almost ready to complete the Module Assessment in which you will have the opportunity to apply your understanding of Le Châtelier's principle and some of the other aspects of equilibrium you have learned in Module 7.

Module 7—Principles of Chemical Equilibrium



Module Summary

In Module 7 you considered the following questions:

- What is happening in a system at equilibrium?
- How do scientists predict shifts in the equilibrium of a system?

The study of chemical equilibrium has many applications. In this module you learned that chemical equilibrium is an important process in biological and chemical systems. Knowledge of chemical equilibrium is used in many areas, including environmental sciences, chemistry, pharmacology, sports medicine, the chemical industry, and engineering.

In this module you were introduced to the basic concepts of chemical equilibrium and to many applications in which knowledge of equilibrium helps to better understand the function of these systems. You were introduced to many chemical systems that demonstrate principles like position of equilibrium, equilibrium law, and Le Châtelier's principle.

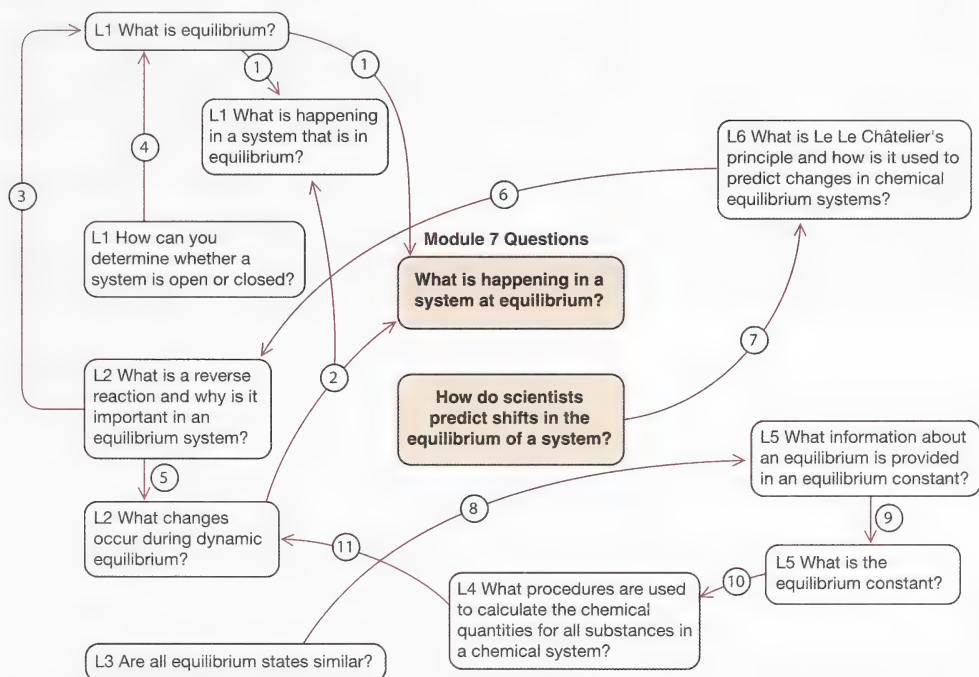
You have kept track of the systems you investigated in a summary table. Now is a good time to use your table to review the concepts you learned by studying the equilibrium systems. Summaries like the table you constructed throughout this module can really help you prepare for upcoming examinations.

You may wish to continue to use a table like the one you used in Module 7 as you investigate acid and base systems and their equilibria in Module 8. The fundamental aspects of chemical equilibrium you learned in Module 7 will be important for your study in Module 8.

Concept Map or Graphic Organizer

As you worked through Module 7 you may have added information to a concept map or graphic organizer, based on the module and lesson questions listed in the Module 7 Concept Organizer. Now is a good time to review the relationships in your concept map or graphic organizer and try to answer the module and lesson questions.

A sample Module 7 concept map is provided below and shows one set of possible links between the questions. Remember that this is one possible description only—there are many other correct possibilities. However, if your completed concept map or graphic organizer differs significantly from the sample, you may wish to contact your teacher or to compare your map or organizer with those of other students in your class. This will ensure that your interpretations of lesson materials, and your descriptions, are accurate.



Linking Statements

#	Statement
1	appears to not change—constant set of macroscopic properties
2	equal rates for the forward and reverse reactions
3	and forward reaction define the changes that occur in an equilibrium
4	Is matter or energy being added/removed?
5	reverse reaction describes one of the ways the system can change
6	stress can favour the reverse or forward reaction to move to a new equilibrium
7	using
8	proportions of products relative to reactants is unique to conditions and described using
9	$[\text{products}]/[\text{reactants}]$
10	is used to
11	ICE tables used to apply stoichiometry to an equilibrium system



Module Assessment

Read the description of the following investigation and view the virtual investigation. Next, complete the Analysis questions. Ensure that your answers include accurate descriptions and explicit reference to the data provided. Remember to state interpretations made from these data using proper representations of, and terminology for, the concepts involved.

Describing a Chemical Equilibrium System

An experiment was performed to investigate the effect of an energy change on the following equilibrium system:



Experimental Design

A colour change in the substances involved will be used to determine whether a change in equilibrium occurs when a test tube containing a mixture of the components in the system is placed in a warm water bath.

Results



Go to the course multimedia DVD, and view the virtual investigation “Describing a Chemical Equilibrium System.”

Analysis

Complete the following questions:

1. Organize your observations from this investigation into a data table.
2. Explain the purpose of test tubes 1, 2, and 3 in the experimental design.
3. Does energy affect the equilibrium system investigated?
4. Write the chemical equation for this equilibrium system, including the term for energy in the equation.
5. What principle was used to investigate the equilibrium in this system? Explain how using this principle allowed you to make interpretations in your answer to question 4.
6. Describe one additional experiment that you could perform to confirm your analysis of this equilibrium system. Include a description of the expected result and an explanation of how the expected result can be used to confirm your analysis.

Save your answers in your course folder and submit a copy to your teacher.

Module 7 Glossary

closed system: a system that is separated from its surroundings by a definite boundary so that energy can enter and leave the system but matter cannot

dynamic equilibrium: the idea that there is a balance between two opposing processes (forward and reverse) occurring at the same rate

The balance between the forward and reverse processes maintains the system's appearance of constant properties, in spite of change that has occurred internally.

equilibrium: the state of a closed system in which the system does not appear to change

For a chemical system, physical properties, like colour, remain constant and concentrations of substances within the system do not change.

macroscopic: can be observed with the unaided eye or other senses

microscopic: cannot be observed with the unaided eye or other senses

open system: a system in which both matter and energy can enter and leave the system

static equilibrium: the idea that forces on an object are opposing and equal

The object remains constant because no changes are happening internally.

Appendix



Module 7 Self-Check Answers

Contact your teacher if your answers vary significantly from the answers provided here.

Lesson 2

SC 1. Answers will vary. Some possible answers are shown below.

Type of Chemical Equilibrium	Example
Phase Equilibrium	humidity in air (gaseous water (vapour) and liquid water) propane cylinder (gaseous propane and liquid propane)
Solubility Equilibrium	concentration of calcium and phosphate ions in blood and as a compound in bones and teeth carbon dioxide gas dissolved into the water in a carbonated beverage dissolved oxygen in a body of water
Chemical Reaction Equilibrium	ionization of an acid or a base (reaction with water to produce hydronium ions (acid) or hydroxide (base))

Lesson 3

SC 1.

Practice 3.

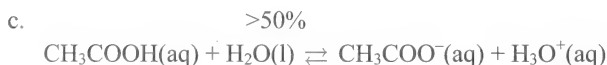
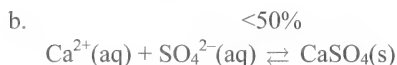
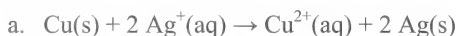
- A system at equilibrium appears not to change, or has a constant set of observable properties.
- An equilibrium involves a forward and a reverse reaction occurring at equal rates. These two reactions involve change at the microscopic level; because change occurs, the system is dynamic.
- The rates of the forward and reverse reactions are equal.

Practice 4.

The maximum possible yield of $\text{CH}_3\text{Cl}(\text{g})$ is 2.00 mol. This is because 2.00 mol of $\text{CH}_4(\text{g})$ is used and it is the limiting reagent, and all reactants combine in 1:1 molar proportions.

b.
$$\begin{aligned} \% \text{ reaction} &= \left(\frac{1.40 \text{ mol}}{2.00 \text{ mol}} \right) 100 \\ &= 70.0\% \end{aligned}$$

The position of the equilibrium favours the products.

Practice 7.**Lesson 4**

SC 1. Mole ratios can be used to complete the calculation for quantity of $\text{CH}_4(\text{g})$ reacted.

$$\begin{aligned} \text{mol}_{\text{methane reacted}} &= \frac{1.40 \text{ mol}_{\text{chloromethane}} \times 1 \text{ mol}_{\text{methane}}}{1 \text{ mol}_{\text{chloromethane}}} \\ &= 1.40 \text{ mol}_{\text{methane}} \end{aligned}$$

The moles of methane that reacted were 1.40 mol.

The quantity of methane remaining is $2.00 \text{ mol} - 1.40 \text{ mol} = 0.60 \text{ mol}$.

SC 2. Using mole ratios will allow the calculation of the remaining substances.

Moles of chlorine reacted will be equal to the moles of methane that reacted, since they combine in 1:1 proportions. Therefore, moles of $\text{Cl}_2(\text{g})$ reacted = 1.40 mol and moles of $\text{Cl}_2(\text{g})$ remaining is $10.00 \text{ mol} - 1.40 \text{ mol} = 8.60 \text{ mol}$.

Moles of $\text{HCl}(\text{g})$ produced will be equal to moles of chloromethane produced, since the reactants have 1:1 proportions (as shown in the balanced chemical equation). Therefore, moles of $\text{HCl}(\text{g})$ = 1.40 mol.

SC 3. I = initial, C = change, and E = equilibrium concentrations of reactants and products in the equilibrium system.

SC 4. The change in concentration of each species is an important value to have because it identifies the quantity of substance that has undergone chemical change as equilibrium has been established. In a system at equilibrium, some quantity of each species in the system will be present; therefore, the change in quantity for one substance is required to interpret the change in quantity for other substances in the system.

SC 5. A negative sign represents a decrease in quantity for that substance. A decrease in chemical quantity or moles for reactants is expected since reactants are consumed by the forward reaction as the system establishes equilibrium.

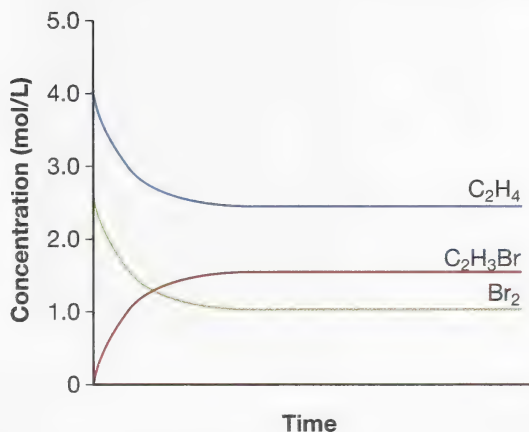
SC 6.

Practice 6.

- a. The graph shows that the equilibrium concentration of ethene is 2.5 mol/L. The change in ethene concentration is $4.0 \text{ mol} - 2.5 \text{ mol} = 1.5 \text{ mol}$. Since the coefficients in the balanced chemical equation are all 1, all species undergo a change in concentration of 1.5 mol/L. Therefore, the equilibrium concentrations are as follows:

$$\text{Br}_2(\text{g}) = 2.50 \text{ mol/L} - 1.5 \text{ mol/L} = 1.0 \text{ mol/L}$$

$$\text{C}_2\text{H}_3\text{Br}(\text{g}) = 0 \text{ mol/L} + 1.50 \text{ mol/L} = 1.50 \text{ mol/L}$$



b.

Concentration	$\text{C}_2\text{H}_6(\text{g}) (\text{mol/L})$	$\text{Br}_2(\text{g}) (\text{mol/L})$	$\text{C}_2\text{H}_5\text{Br}(\text{g}) (\text{mol/L})$
Initial	4.0	2.5	0
Change	-1.5	-1.5	+1.5
Equilibrium	2.5	1.0	1.5

- c. The volume of the container is 1 litre. This is determined by the information provided on the graph, which shows an initial concentration of ethane of 4.0 mol/L. Since 4.0 mol of ethane were used, the volume of the container must be 1.0 L.
- d. Since all coefficients are 1, the moles of bromine are lower than the moles of ethene. Therefore, bromine is the limiting reagent. The maximum yield of bromoethene would be 2.50 mol. Therefore,

$$\begin{aligned}\% \text{ yield} &= \left(\frac{1.5 \text{ mol}}{2.5 \text{ mol}} \right) 100 \\ &= 60\%\end{aligned}$$

SC 7.

Concentration	$\text{N}_2(\text{g}) (\text{mol/L})$	$\text{H}_2(\text{g}) (\text{mol/L})$	$\text{NH}_3(\text{g}) (\text{mol/L})$
Initial	$\frac{2.42 \text{ mol}}{2.00 \text{ L}} = 1.21 \text{ L}$	$\frac{21.6 \text{ mol}}{2.00 \text{ L}} = 10.8$	0
Change	$1.05 \times \left(\frac{1}{2} \right) = 0.53$		$0 + 1.05 = +1.05$
Equilibrium	$1.21 - 0.53 = 0.68$	$10.8 - 1.58 = 9.2$	1.05

Lesson 5

SC 1.

Section 15.1

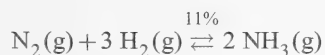
1.a.	Equation	$\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightleftharpoons 2 \text{HCl}(\text{g})$
	Equilibrium Law Expression	$K_c = \frac{[\text{HCl}(\text{g})]^2}{[\text{H}_2(\text{g})][\text{Cl}_2(\text{g})]}$
1.b.	Equation	$\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightleftharpoons 2 \text{NH}_3(\text{g})$
	Equilibrium Law Expression	$K_c = \frac{[\text{NH}_3(\text{g})]^2}{[\text{N}_2(\text{g})][\text{H}_2(\text{g})]^3}$
1.e.	Equation	$\text{CaCO}_3(\text{s}) \rightleftharpoons \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$
	Equilibrium Law Expression	$K_c = \frac{\cancel{[\text{CaCO}_3(\text{s})]}^1 [\text{CO}_2(\text{g})]^2}{[\cancel{\text{CaCO}_3(\text{s})}]^1}$ $= [\text{CO}_2(\text{g})]$
1.f.	Equation	$\text{Na}_2\text{SO}_4(\text{aq}) + \text{CaCl}_2(\text{aq}) \rightleftharpoons 2 \text{NaCl}(\text{aq}) + \text{CaSO}_4(\text{s})$ Net ionic equation: $\text{SO}_4^{2-}(\text{aq}) + \text{Ca}^{2+}(\text{aq}) \rightleftharpoons \text{CaSO}_4(\text{s})$
	Equilibrium Law Expression	$K_c = [\text{Ca}^{2+}(\text{aq})][\text{SO}_4^{2-}(\text{aq})]$

Section 15.1 2.

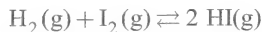
Alkenes tend to readily undergo addition reactions with halogens. It would be expected that the reaction would be stoichiometric, or greater than 99% complete.

Section 15.1 4.

a.



b.

**Section 15.1 5.**

$$K_c = \frac{[\text{HI(g)}]^2}{[\text{H}_2\text{(g)}][\text{I}_2\text{(g)}]}$$

$$\begin{aligned} \text{System 1} \quad K_c &= \frac{[7.80 \text{ mmol/L}]^2}{[1.10 \text{ mmol/L}][1.10 \text{ mmol/L}]} \\ &= 50.3 \end{aligned}$$

$$\begin{aligned} \text{System 2} \quad K_c &= \frac{[2.10 \text{ mmol/L}]^2}{[0.30 \text{ mmol/L}][0.30 \text{ mmol/L}]} \\ &= 49 \end{aligned}$$

$$\begin{aligned} \text{System 3} \quad K_c &= \frac{[2.5 \text{ mmol/L}]^2}{[0.35 \text{ mmol/L}][0.35 \text{ mmol/L}]} \\ &= 51 \end{aligned}$$

Section 15.1 6.

$$\begin{aligned} K_c &= \frac{[\text{NH}_3\text{(g)}]^2}{[\text{N}_2\text{(g)}][\text{H}_2\text{(g)}]^3} \\ 6.0 \times 10^{-2} &= \frac{[\text{NH}_3\text{(g)}]^2}{[0.50 \text{ mol/L}][1.50 \text{ mol/L}]^3} \\ [\text{NH}_3\text{(g)}] &= \sqrt{6.0 \times 10^{-2} (0.50 \text{ mol/L})(1.50 \text{ mol/L})^3} \\ &= 0.32 \text{ mol/L} \end{aligned}$$

Section 15.1 7.

a.



b.

$$K_c = \frac{[\text{H}_2(\text{g})][\text{Br}_2(\text{g})]}{[\text{HBr}(\text{g})]^2}$$

c.

$$\begin{aligned}\text{Moles HBr at equilibrium} &= 0.100 \text{ mol/L} \times 2.00 \text{ L} \\ &= 0.200 \text{ mol}\end{aligned}$$

d.

$$\begin{aligned}\text{Change in amount of HBr} &= 1.00 \text{ mol} - 0.200 \text{ mol} \\ &= 0.800 \text{ mol}\end{aligned}$$

e.

$$\begin{aligned}\text{Change in H}_2(\text{g}) &= 0.800 \text{ mol} \times \left(\frac{1 \text{ H}_2}{2 \text{ HBr}} \right) \\ &= 0.400 \text{ mol}\end{aligned}$$

Since H_2 and Br_2 are produced in 1:1 proportions, then the chemical amount of Br_2 formed is 0.400 mol.

f. Concentration at equilibrium:

$$\text{HBr}(\text{g}) = 0.100 \text{ mol/L}$$

$$\begin{aligned}\text{H}_2(\text{g}) &= \frac{0.400 \text{ mol}}{2.00 \text{ L}} \\ &= 0.200 \text{ mol/L}\end{aligned}$$

$$\begin{aligned}\text{Br}_2(\text{g}) &= \frac{0.400 \text{ mol}}{2.00 \text{ L}} \\ &= 0.200 \text{ mol/L}\end{aligned}$$

g.

$$\begin{aligned}K_c &= \frac{[0.200 \text{ mol/L}][0.200 \text{ mol/L}]}{[0.100 \text{ mol/L}]^2} \\ &= 4.00\end{aligned}$$

Section 15.1 9.

Assume the flask contains no other substances at the start of the experiment except CO(g) and $\text{H}_2\text{O(g)}$ and that the reaction chamber has a volume of 1.00 L.

Concentration	$\text{Cl}_2\text{(g)}$ (mol/L)	$\text{H}_2\text{(g)}$ (mol/L)	CO(g) (mol/L)	$\text{H}_2\text{O(g)}$ (mol/L)
Initial	0	0	0.20	0.25
Change	+0.10	+0.10	-0.10	-0.10
Equilibrium	0.10	0.10	0.10	0.15

$$\begin{aligned}
 K_c &= \frac{[\text{CO(g)}][\text{H}_2\text{O(g)}]}{[\text{CO}_2\text{(g)}][\text{H}_2\text{(g)}]} \\
 &= \frac{[0.10 \text{ mol/L}][0.15 \text{ mol/L}]}{[0.10 \text{ mol/L}][0.10 \text{ mol/L}]} \\
 &= 1.5
 \end{aligned}$$

Section 15.1 10.

Assume the flask contains no other substances at the start of the experiment except $\text{H}_2\text{(g)}$ and $\text{Br}_2\text{(g)}$.

Let x = change in concentration of hydrogen

Concentration	2 HBr(g) (mol/L)	$\text{H}_2\text{(g)}$ (mol/L)	$\text{Br}_2\text{(g)}$ (mol/L)
Initial	0	0.50	0.50
Change	+2x	-x	-x
Equilibrium	2x	0.50 - x	0.50 - x

$$\begin{aligned}
 K_c &= \frac{[\text{H}_2(\text{g})][\text{Br}_2(\text{g})]}{[\text{HBr}(\text{g})]^2} \\
 0.020 &= \frac{[0.50 - x][0.50 - x]}{[2x]^2} \\
 \sqrt{0.020} &= \sqrt{\frac{[0.50 - x]^2}{[2x]^2}} \\
 0.14 &= \frac{[0.50 - x]}{2x} \\
 0.28x &= 0.50 - x \\
 1.28x &= 0.50 \\
 x &= \frac{0.50}{1.28} \\
 &= 0.39
 \end{aligned}$$

Therefore, the change in concentration of $\text{H}_2(\text{g})$ is 0.39 mol/L.

$$\begin{aligned}
 n_{\text{H}_2(\text{g})} &= n_{\text{Br}_2(\text{g})} \\
 &= 0.11 \text{ mol/L} \times 0.500 \text{ L} \\
 &= 0.055 \text{ mol}
 \end{aligned}$$

$$\begin{aligned}
 n_{\text{HBr}(\text{g})} &= 0.78 \text{ mol/L} \times 0.500 \text{ L} \\
 &= 0.38 \text{ mol}
 \end{aligned}$$

The equilibrium concentrations of all substances are

$$\begin{aligned}
 \text{HBr}(\text{g}) &= 2x \\
 &= 2(0.39 \text{ mol/L}) \\
 &= 0.78 \text{ mol/L}
 \end{aligned}$$

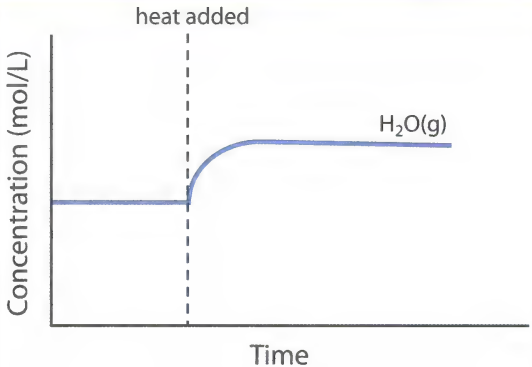
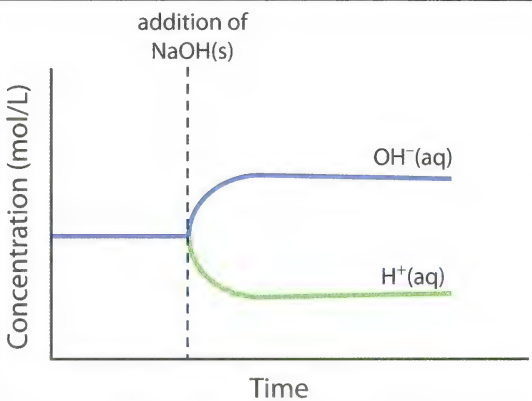
$$\begin{aligned}
 \text{H}_2(\text{g}) &= 0.50 - x \\
 &= 0.50 - 0.39 \\
 &= 0.11 \text{ mol/L}
 \end{aligned}$$

$$\begin{aligned}
 \text{Br}_2(\text{g}) &= 0.50 - x \\
 &= 0.50 - 0.39 \\
 &= 0.11 \text{ mol/L}
 \end{aligned}$$

Lesson 6

SC 1.

Practice 2.

	Direction of Shift in Equilibrium	Graph
A	Right	 Concentration (mol/L) Time heat added $\text{H}_2\text{O(g)}$
B	Left	 Concentration (mol/L) Time addition of NaOH(s) $\text{OH}^{\text{-}}(\text{aq})$ $\text{H}^+(\text{aq})$

Practice 3.

- Increasing the pressure increases the frequency of collisions between the gaseous particles favouring the forward reaction. The system shifts to the right.
- The forward reaction is exothermic; therefore, increasing the temperature will favour the reverse reaction. The system shifts to the left.
- A high temperature is necessary to speed up the reactions so the process occurs in a reasonable amount of time.
- Both actions, increasing the concentration of reactants and removing the product, are stresses that will favour the forward reaction in the system.

D	Right	
E	Right	

Chemistry 30

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Acid-Base Equilibrium

Module 8

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Module 8—Acid-Base Equilibrium

Module Introduction

In Module 7 you learned about chemical equilibrium. In this module you will apply your understanding of the principles of equilibrium to chemical systems containing aqueous acids and bases. You will use your knowledge of dynamic equilibrium, equilibrium position, and shifts in equilibrium. You will also rely on your ability to analyze an equilibrium system using quantitative techniques (like calculating a value for the equilibrium constant, K_c) and your ability to complete an ICE table.

Your study of the equilibrium of aqueous acids and bases will include an investigation of environmental and biological systems. You will investigate how the components of these systems can interact to form equilibrium systems, and you will consider how the behaviour of these systems can be better understood through knowledge of the principles of chemical equilibrium.



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In Module 8 you will investigate the following question:

- How does the equilibrium of acids and bases affect environmental and biological systems?

In Modules 5 and 6 you used a Table of Selected Standard Electrode Potentials to analyze redox systems. The Module 8 Assessment will require you to consider the construction of the Table of Relative Strengths of Acids and Bases, which you will use throughout this module.

Remember that each lesson will also be organized around questions intended to guide your study. As you proceed through Module 8, you may record answers to these questions and any interrelationships that exist between them in a concept map or graphic organizer. More information is available in the Unit D Concept Organizer. In the Module 8 Summary you will receive further information on how you can use your concept map or graphic organizer to review the concepts you studied in this module.



Big Picture

Soil is often overlooked when we think of important parts of Earth. In earlier science courses you learned that soil is a component of the biosphere, the section of Earth that supports life. Apart from being the medium for growing most plants, soil is the home for many organisms. Soil is also made up of many chemical components. The minerals, organic materials, and water within soil form a complex chemical system.



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The atmosphere also forms a system in which equilibria can exist. How might these equilibria affect the rain and other forms of precipitation that will deposit on Earth?

In Module 7 you learned that certain combinations of chemicals can form an equilibrium system. You also learned that equilibrium systems exhibit certain unique properties that can be manipulated by the application of a stress.

What kinds of stresses are placed on soil? Does human activity or agricultural practice affect the equilibrium of chemical components in soil? How would the chemical components in soil respond to the application of a stress, and what impact might that stress have on soil's fertility?

Soil is not the only system that has chemical components that form an equilibrium system. In this module you will also investigate the components within biological systems that make up equilibrium systems within organisms.



Assessment in This Module

Each lesson contains a range of activities and assessment options. These include assignments, labs, and Self-Check, Try This, Discuss, Reflect and Connect, and Reflect on the Big Picture activities. Instructions will be provided for each of these activities so that you can appropriately focus your time and effort. Your teacher will tell you which assessment options to complete and which responses to submit for marks or feedback. Remember to save all of your work in your Chemistry 30 folder.

In the Module 8 Assessment you will consider the construction of the Table of Relative Strengths of Acids and Bases, which you will use throughout this module.

You may wish to look at the Module Assessment and the Unit Assessment before starting Lesson 1.

In This Module

Lesson 1—Introduction to the Equilibrium of Acids and Bases

In Lesson 1 you will be introduced to the equilibrium of acids and bases and the equilibrium of water. You will revisit the concepts of the hydronium ion and of the acidic particle. You will also recall how hydronium ions are formed in aqueous solutions and are believed to be responsible for the properties of acidic solutions.

You will investigate the following lesson questions:

- What is the hydronium ion and how does it account for the properties of acidic solutions?
- What chemical components are part of the equilibrium of water?
- How does the equilibrium of water affect the calculation of pH?

Lesson 2—Acid Strength and Equilibrium

Acids can be classified as either strong acids or weak acids. In Lesson 2 you will apply your understanding of equilibrium to investigate the strength of acids.

You will investigate the following lesson question:

- How is knowledge of equilibrium used to explain the differences in acid strength?

Lesson 3—Brønsted-Lowry Theory

In Lesson 3 you will learn about the development of the Brønsted-Lowry theory. The Brønsted-Lowry theory explains the chemical behaviour of acids and bases.

You will investigate the following lesson questions:

- What occurs during a chemical reaction between an acid and a base?
- How does the Brønsted-Lowry theory support what is known about the equilibrium of aqueous acids and bases?
- How do conjugate acid-base pairs represent an equilibrium system?

Lesson 4—Predicting AB Equilibrium

In Lesson 4 you will test a method to predict the equilibrium position for acid-base reactions.

You will investigate the following lesson question:

- How is knowledge of the equilibrium of aqueous systems involved in predicting the outcome of acid-base reactions?

Lesson 5—Equilibrium Law and the Strength of Acids and Bases

In Lesson 5 you will investigate the use of quantitative techniques to describe the equilibrium of aqueous acids and bases.

You will investigate the following lesson questions:

- What are K_a , K_b , and K_w ?
- How do K_a and K_b explain the position of the equilibrium of aqueous acids and bases?
- How are values for K_a and K_b used to calculate the pH of solutions containing weak acids and bases?

Lesson 6—pH Curves

In Lesson 6 you will review acid-base titrations, and you will investigate the parts of a pH curve that demonstrate behaviour as equilibrium systems. You will also learn about acidic and basic substances that can react multiple times and how this process appears on a pH curve.

You will investigate the following lesson questions:

- What information about acids and bases and their equilibrium is contained on a titration curve?
- What events occur in the reaction of polyprotic acids and bases?

Lesson 7—Buffers

In Lesson 7 you will investigate the properties of buffers as equilibrium systems.

You will investigate the following lesson questions:

- What is a buffer and how does a buffer exhibit the characteristics of an equilibrium system?
- What is buffering capacity?

Module 3 – Acid-Base Equilibrium

Lesson 1—Introduction to the Equilibrium of Acids and Bases



Get Focused

Successful farming is not just a matter of planting a crop and seeing what grows. Farming operations cannot control the weather, one key component for success, but they can improve their success by carefully managing the soil.



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In the previous module you learned about chemical equilibrium

and how chemical systems at equilibrium can be stressed. In this module you will learn about some of the chemical components of soil that are involved in an equilibrium system, and you will discover how some aspects of soil management focus on this equilibrium system.

One aspect of soil management is salinity, which is caused by the concentration of chloride ions in the soil. Soils that experience high saline content can appear to have abundant water, but they cannot support plant growth. In this case, the poor plant growth is due to the effect chloride ions have on the ability of plants to absorb available water. You might think that absorbing water is a simple process, but the movement of water and water's equilibrium in the cells of plant roots and in the soil can be affected by the concentration of chloride ions.

In this lesson you will learn about the equilibrium that exists within water and how this equilibrium is influenced by the presence of acids and bases within an aqueous system.

Consider the following questions as you complete Lesson 1:

- What is the hydronium ion and how does it account for the properties of acidic solutions?
- What chemical components are part of the equilibrium of water?
- How does the equilibrium of water affect the calculation of pH?



Module 8: Lesson 1 Assignment

There is no assignment for this lesson.

There are other questions in this lesson that are not marked by the teacher; however, you should still answer these questions. The Self-Check, Try This, and other types of questions are placed in this lesson to help you review important information and build key concepts that may be applied in future lessons. You should record the answers to all the questions in the lesson and place those answers in your course folder.

After a discussion with your teacher, you must decide what to do with the questions that are not part of your assignment. For example, you may decide to submit the responses to Try This and other questions that are not marked to your teacher for informal assessment and feedback. Your answers are very important to your teacher. They provide your teacher with information about your learning, and they help your teacher identify where adjustments to your instruction may be necessary.



Explore



Try This

Summarize What You Know

This module builds on your knowledge of acids and bases gained from previous chemistry courses. You may wish to review the following parts of the textbook to prepare for the remainder of your study in this module:

- “The Nature of Acidic and Basic Solutions,” page 236
- various content about the hydronium ion, page 237
- “pH and pOH Calculations,” pages 238–244
- “Acid–Base Indicators,” page 245
- “The Strength of Acids and Bases,” pages 254–257
- “Explaining Acids and Bases,” pages 248–251
- “Polyprotic Substances,” page 258
- “Titration Analysis,” pages 328–330
- “Acid–Base Titration Curves and Indicators,” pages 333–336

Not in the mood for all that reading? Try entering the word “video” with each of the terms in the bullet list in an Internet search.

Regardless of how you choose to review these concepts, keep a summary of relevant reactions and concepts for each topic. Your summary for each topic might take the form of a list, outline, concept map, or other form. Remember to add to your summaries as you work through this module.



Read

In Chemistry 20 you learned about the existence of the hydronium ion, $\text{H}_3\text{O}^+(\text{aq})$, the acidic particle in aqueous solutions. The hydronium ion is produced by a chemical reaction in which a water molecule becomes protonated. What process allows for a water particle to gain a proton?

You may have hypothesized what would happen if two water molecules were to collide. The resulting reaction would create species that would become part of an equilibrium system within water.

Read pages 713–716 and work through “Communication examples” 1–4 of the textbook to learn more about the equilibrium of water and its equilibrium constant, K_w .



Self-Check

SC 1. Complete “Practice” questions 1–4 on page 716 of the textbook.

Check your work with the answer in the appendix.



Discuss

In the previous section you used the water ionization constant, K_w , to perform calculations. In Module 7 you learned to interpret the value of equilibrium constants like K_w .

D 1. What does the value for K_w mean about the position of the equilibrium of water?

D 2. Which reaction is favoured in this equilibrium, the forward or reverse, and to what degree?

D 3. What empirical observations do you have to support your answers to D 1 and D 2?

Post your answers to the discussion area for your class, and submit a copy to your teacher.



Read

Your ability to calculate solution pH and pOH is critical to your investigation of the equilibrium of acids and bases. In Chemistry 20 you were restricted in your ability to perform calculations by not knowing how to convert between concentrations of hydroxide ions and hydronium ions within a solution. With your understanding of the equilibrium of water from the previous section, you can now use K_w to quickly complete such calculations.

Read “Communicating Concentrations: pH and pOH” on pages 716–717 of the textbook. Remember to add new information to the appropriate summary you began at the start of this lesson.



Self-Check

SC 2. Complete “Practice” questions 7–8 on page 718 of the textbook.

Check your work with the answer in the appendix.



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Reflect and Connect

Earlier in this lesson you learned that the salinity of soil affects plants’ ability to absorb water. In soils with high salinity, the large concentration of chloride ions in the soil retains water, preventing the water from being absorbed by plants.

Soil salinity is the result of the leaching of chloride ions from salt formations beneath the surface and accumulations of the solutes where water collects and evaporates. In areas with high salinity, like the one shown in the photograph, a solubility equilibrium exists where salt precipitates and dissolves depending on the availability of water.



Complete the following investigation to analyze the solubility of chloride ions in the equilibrium provided.

Read “Exploration: Salty Acid or Acidic Salt?” on page 711 of the textbook. Then go to the course multimedia DVD, and view “Exploration: Salty Acid or Acidic Salt?”



Self-Check

SC 3. Complete “Exploration: Salty Acid or Acidic Salt?” questions a–f on page 711 of the textbook.

Check your work with the answer in the appendix.



Reflect and Connect

You initiated summaries earlier in this lesson. Update your summaries to include concepts you learned in this lesson. Save a copy of your updated summaries in your course folder. You may be asked to submit a copy of your summaries to your teacher or to share one or more of your summaries with other students in your class.



Lesson Summary

In Lesson 1 you considered the following questions:

- What is the hydronium ion and how does it account for the properties of acidic solutions?
- What chemical components are part of the equilibrium of water?
- How does the equilibrium of water affect the calculation of pH?

You learned that pure water is a chemical system with an equilibrium between hydronium ions and hydroxide ions, albeit at extremely low concentrations. You used the water ionization constant to calculate the molar concentrations of $[\text{H}_3\text{O}^+(\text{aq})]$ and $[\text{OH}^-(\text{aq})]$ that exist within water as a result of water’s equilibrium. You also reviewed the calculation of pH and pOH, further incorporating your knowledge of the equilibrium of water.

Module 8—Acid-Base Equilibrium

Lesson 2—Acid Strength and Equilibrium



Get Focused

Acid rain is an environmental issue that has received much attention. However, rain is only one mechanism by which acidic material from the atmosphere can meet Earth's surface. Acid deposition is a broader term that refers to all forms of wet and dry deposits that have acidic properties. How can rain, snow, and even dust become acidic? Can you recall the chemical reaction that produces the acidic particle, the hydronium ion, and explains the properties of acids?

Four acids have been implicated in acid deposition. The four acids include nitrous and nitric acids, from emissions of nitrous oxides, and sulfurous and sulfuric acids, from emissions of sulfur oxides. Are there any differences in properties among these four acids? Which of these acids poses the greatest risk to the environment?



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In your previous chemistry course you learned about the concept of weak and strong acids. How does this classification relate to the chemical reaction described above and its equilibrium? Which of the four acids are categorized as strong? Which are categorized as weak? How might the classification of these substances explain their effects on the environment?

Consider the following question as you complete Lesson 2:

- How is knowledge of equilibrium used to explain the differences in acid strength?



Module 8: Lesson 2 Assignment

You will be directed to go to the Assignment Booklet to complete the Module 8: Lesson 2 Assignment later in this lesson.

You must decide what to do with the questions that are not marked by the teacher.

Remember that these questions provide you with the practice and feedback that you need to successfully complete this course. You should respond to all the questions and place those answers in your course folder.



Explore



Read

In Module 7 you used information about the extent of reaction to understand equilibrium position. Another way to understand equilibrium position is to observe the properties of a system. When considering acids, you learned that strong acids react (ionize) 100% with water, whereas weak acids react (ionize) less than 100% with water. Which properties of an acidic solution would be affected by the differences in the extent of reaction (ionization)?

Read “Acid Strength as an Equilibrium Position” on pages 718–720 of the textbook. Work through “Communication example 5” and “Communication example 6” on page 720 to learn how percent ionization is used in the calculation of pH and how a pH value can be used to determine a percent ionization for an acid.



Self-Check

Complete “Section 16.1” questions 1, 2, 4, 6, and 7 on page 721 of the textbook.

Check your work with the answer in the appendix.



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Discuss

In the previous Read section you learned that the distinction between a strong and a weak acid is the extent to which the acid ionizes. How does this definition fit with your understanding of an equilibrium system?

Hydrochloric acid is a strong acid, whereas hydrofluoric acid is a weak acid. Use a balanced chemical equation to describe the equilibrium position of these two different types of acids.

Post a copy of your response to the discussion area for your class. Save a copy of your response in your course folder and submit a copy to your teacher.



Reflect on the Big Picture

Think back to the Module 8 Big Picture, which discussed examples of equilibrium in the soil, water, and atmosphere.

The combustion process that occurs in automobile engines and furnaces produces nitrous oxides, $\text{NO}(\text{g})$ and $\text{NO}_2(\text{g})$. The following unbalanced series of reactions describes the deposition of nitric acid as a result of the chemical joining of nitrogen and oxygen that occurs during a combustion reaction.



Use your knowledge of equilibrium to explain how the energy from high-temperature combustion processes contributes to the production of $\text{HNO}_3(\text{g})$.

Save a copy of your response in your course folder. You will further evaluate your answer later in this module.



Reflect and Connect

In Lesson 1 you initiated a summary about the strength of acids and bases. Take this opportunity to update your summary to include concepts you learned in this lesson. Save a copy of your updated summary in your course folder. You may be asked to submit a copy of this summary to your teacher for feedback.



Module 8: Lesson 2 Assignment

Go to the Assignment Booklet, and complete the “Module 8: Lesson 2 Assignment.”



Lesson Summary

In Lesson 2 you considered the following question:

- How is knowledge of equilibrium used to explain the differences in acid strength?

Acids can be classified as either strong or weak as a result of their ionization, which produces hydronium ions. Strong acids ionize completely. You can consider the concentration of hydronium ions that results from a strong acid's ionization to be equal to the initial concentration of the acid present in the solution. Weak acids ionize to a small extent. In the case of weak acids, the extent of ionization must be determined or known in order to predict the resulting concentration of hydronium ions.

Knowledge of equilibrium allows for a description of the ionization of strong and weak acids and for a calculation of the acid's percentage ionization. In Lesson 5 you will further study the equilibrium of strong and weak acids using the equilibrium lab and the ionization constant for acids.

Module 8—Acid-Base Equilibrium

Lesson 3—Brønsted-Lowry Theory



Get Focused

The concept of acids you currently work with involves a reaction with water to produce hydronium ions. As you know from earlier in this module, bases dissociate to produce hydroxide ions. Those descriptions work well for substances like hydrochloric acid and



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calcium hydroxide, but how do you explain the acidic properties of the ammonium ion and the basic properties of the carbonate ion?

Ammonium is commonly used in nitrogen-based fertilizers, and carbonate is one of many chemical components in soil—especially soils in Alberta. What is necessary to better understand the acidic and basic properties of these substances? Would this understanding help you to predict changes that might occur to soil as a result of soil management?

Consider the following questions as you complete Lesson 3:

- What occurs during a chemical reaction between an acid and a base?
- How does the Brønsted-Lowry theory support what is known about the equilibrium of aqueous acids and bases?
- How do conjugate acid-base pairs form an equilibrium system?



Module 8: Lesson 3 Assignment

There is no assignment for this lesson.

You must decide what to do with the questions that are not marked by the teacher.

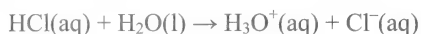
Remember that these questions provide you with the practice and feedback that you need to successfully complete this course. You should respond to all the questions and place those answers in your course folder.



Explore

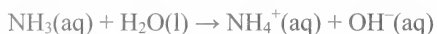
At least one of the summaries you started in Lesson 1 will have covered early theories for acids and bases. The early theories proposed that dissociation of solutes to produce hydrogen ions or hydroxide ions could explain the acidic and basic properties of solutions. A flaw of these theories was the lack of an ability to explain the properties of common substances that clearly have acidic or basic properties.

A modification to these early theories required the consideration of the possibility of a chemical reaction between the solute and water. For example, the reaction between aqueous hydrogen chloride, HCl(aq) , with water is



The product of this reaction is the hydronium ion; it was later confirmed to be the acidic particle.

The reaction of a substance like ammonia, NH_3 , with water is



The earlier theory could not explain the basic properties of aqueous ammonia, but the reaction with water shows the probability of hydroxide ions being produced as products of the reaction. Simple diagnostic tests, like litmus, confirm this theoretical explanation.



Discuss

In previous lessons you considered each reaction as an equilibrium, and you used methods to communicate equilibrium position. Consider the following systems:

System	$\text{HCl(aq)} + \text{H}_2\text{O(l)} \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{Cl}^-(\text{aq})$
Concentration of HCl(aq)	0.0250 mol/L
Solution pH	1.6

System	$\text{NH}_3(\text{aq}) + \text{H}_2\text{O(l)} \rightarrow \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$
Concentration of $\text{NH}_3(\text{aq})$	0.0250 mol/L
Solution pH	10.8

D 1. Explain how you would use the information provided to determine the equilibrium positions for the two systems. Give reasons for your suggested approach.

Post your proposed strategy to the discussion area for your class. Read the strategies of at least two other students. How are the other students using the data provided? Will their strategies work? Provide constructive feedback to your classmates regarding the strategies they posted. Revise your strategy based on your readings and the feedback you receive. Make sure you record the reason for any modifications made. Save your work in your course folder.

D 2. Use the information provided to determine and appropriately communicate the equilibrium position for the reactions shown.

Save a copy of your answers in your course folder and submit a copy to your teacher.



Read

You might conclude that it appears that many substances require a reaction with water to exhibit their properties. Do any general trends exist in these reactions?

Read pages 722–724 in the textbook to examine the trends that might exist in the reactions of acids and bases.

Brønsted-Lowry Concept

In the textbook reading you learned that the Brønsted-Lowry concept redefines how you will consider acids and bases as entities in a chemical reaction. According to this theory you will describe acids and bases as a result of the kind of chemical change they undergo, rather than by their empirical properties.

Because you have seen that some substances have the ability to react as either an acid or a base, you will further modify your concept of acids and bases. Later in this module you will learn how to predict when these substances will act in one of these possible roles.



Self-Check

SC 1. Complete “Practice” questions 1–5 on page 724 of the textbook.

Check your work with the answer in the appendix.



Read

The proton transfer reaction during acid-base reactions, as described by the Brønsted-Lowry concept, is unique. For instance, the reaction of ammonia, $\text{NH}_3(\text{aq})$, produces the ammonium ion, $\text{NH}_4^+(\text{aq})$. Note that the difference between these two chemical formulas is one proton. More interesting is that it is foreseeable that the product could easily donate a proton in a future reaction, thus acting as a Brønsted-Lowry acid.

Read “Conjugate Acids and Bases” on pages 724–726 in the textbook to learn more about acid-base pairs.



Self-Check

SC 2. Complete “Practice” questions 7–8 on page 726 of the textbook.

Check your work with the answer in the appendix.



Try This

In the previous section you learned about the proton transfer concept. In this activity you will further test this theory by applying material learned in this and previous lessons.

TR 1. The following acids are responsible for acid deposition. Can their acidity be explained using the proton transfer concept?

- a. $\text{HNO}_3(\text{aq})$
- b. $\text{HNO}_2(\text{aq})$
- c. $\text{H}_2\text{SO}_4(\text{aq})$
- d. $\text{H}_2\text{SO}_3(\text{aq})$

TR 2. Carbonate ions, CO_3^{2-} , from limestone, the mineral calcium carbonate, are a component of soils in Alberta. Soils in Alberta tend to be alkaline, as does the pH of the water in most Alberta lakes. Is it possible that the observed pH of soil and lake water could be due to the presence of carbonate in soil?

TR 3. Acid deposition has had relatively little effect on the pH of soils in most parts of Alberta. Use a chemical reaction to explain this observation. In the reaction you write, identify the conjugate acid-base pairs.

Submit your answers to your teacher for feedback.



Reflect and Connect

In Lesson 1 you started summaries about your previous learning on acids and bases. Take this opportunity to update the appropriate summary to include concepts you learned in this lesson. Save a copy of your updated summary in your course folder. You may be asked to submit a copy to your teacher for feedback.



Lesson Summary

In Lesson 3 you considered the following questions:

- What occurs during a chemical reaction between an acid and a base?
- How does the Brønsted-Lowry theory support what is known about the equilibrium of aqueous acids and bases?
- How do conjugate acid-base pairs form an equilibrium system?

The Brønsted-Lowry concept states that acids act as proton donors and bases act as proton acceptors during chemical reactions. This concept explains the observations of acidic or basic properties of substances that previous theories could not explain.

Further empirical evidence to support this concept comes from the changes to pH that occur during some chemical reactions. These changes can only be explained if it is assumed that a proton transfer has occurred between reacting species.

Finally, you learned to consider a Brønsted-Lowry reaction in terms of conjugate pairs of acids and bases, and you learned about the potential of these systems to act in equilibrium.

Module 8—Acid-Base Equilibrium**Lesson 4—Predicting AB Equilibrium****Get Focused**

In Lesson 3 you learned about the Brønsted-Lowry concept. You were able to use this principle to explain the chemical change occurring in an acid-base reaction. However, this principle is not able to explain why certain acid-base combinations will react and others will not. In this lesson



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you will investigate how

to predict the outcome of acid-base reactions and to determine the equilibrium position for such reactions.

In the previous lesson you investigated the reaction between carbonate ions found in soil and nitric acid, which is responsible for acid deposition. You were able to identify this reaction that might neutralize the nitric acid. However, with your current knowledge, you have no way to know whether the reaction between the two components will occur when the components are mixed.

Consider the following question as you complete Lesson 4:

- How is knowledge of the equilibrium of aqueous systems involved in predicting the outcome of acid-base reactions?

In this lesson you will begin to use the “Relative Strengths of Acids and Bases at 298.15 K” table located on pages 8 and 9 of the Chemistry Data Booklet. You may recall that the Module 8 Assessment requires you to consider the structure of this table. It may be helpful to review the Module 8 Assessment before you start this lesson.

**Module 8: Lesson 4 Assignment**

In this lesson you will complete the lab “Predicting Acid-Base Reactions.” You will be directed to go to the Assignment Booklet to complete the Module 8: Lesson 4 Assignment later in this lesson.

You must decide what to do with the questions that are not marked by the teacher.

Remember that these questions provide you with the practice and feedback that you need to successfully complete this course. You should respond to all of the questions and place those answers in your course folder.



Explore



Read

Read the first paragraph of “Predicting Acid-Base Reaction Equilibria” on page 727 of the textbook.

Empirical data is necessary to develop general trends about the tendency of acids and bases to transfer protons. Review the “Relative Strengths of Acids and Bases at 298.15 K” table located on pages 8 and 9 of the Chemistry Data Booklet. This table is also available on page 829 of the textbook. A few points about the structure of this table in the Chemistry Data Booklet are as follows:

- Columns 1 and 2 list the name and chemical formula for acids. Column 3 contains the chemical formula of the conjugate base for the acid shown in each row.
- Acids are listed by decreasing strength. For example, perchloric acid is a stronger acid than citric acid; as a result, citric acid is listed further down the table.
- The acids listed above the shaded hydronium ion row are strong acids; acids listed below the hydronium ion row are weak acids.
- Bases are listed by increasing strength, with the strongest base being hydroxide. The remaining substances listed in the conjugate base column are weak bases and are listed by increasing strength. For example, the perchlorate ion, $\text{ClO}_4^-(\text{aq})$, is the weakest base on the table. Substances listed below the perchlorate ion demonstrate increasing strength as bases.



Try This

Predicting Acid-Base Reactions

Earlier in this course you used the “Table of Selected Standard Electrode Potentials,” found on page 7 of the Chemistry Data Booklet, to predict and analyze reduction-oxidation reactions. Can the “Relative Strengths of Acids and Bases at 298.15 K” table be used to predict acid-base reactions in a similar way? Read “Lab Exercise 16.B” on page 727 of the textbook to test this possibility.

Acid-Base Equilibrium

In “Lab Exercise 16.B,” examine the positions of the substances involved in reactions 1–6. Note the position of the equilibrium for each of these reactions. Comment on any general trends you see.

Save your comments in your course folder and submit a copy to your teacher for feedback.



Read

Read “Predicting Acid-Base Reactions” on pages 728–731 of the textbook to confirm the trends you may have observed in “Lab Exercise 16.B.”

Predicting the equilibrium position for acid-base reactions involves identifying reacting species and their positions on the “Relative Strengths of Acids and Bases at 298.15 K” table with greater specificity than when you predicted redox reactions.

To successfully predict acid-base reactions, you must list the reaction substances. Already in this module you have learned that the extent to which strong and weak acids react with water differs. With this in mind, what is the most appropriate way to list the species present in a solution of hydrochloric acid?

The following chart summarizes the way substances should be considered when predicting acid-base reactions and the reason for each guideline:

Listing for Substance	Reason
All strong acids are written as $\text{H}_3\text{O}^+(\text{aq})$.	Strong acids ionize 100% in aqueous solution; therefore, unionized acid particles are not present except at extremely low concentrations.
Strong bases (hydroxides) are listed as $\text{OH}^-(\text{aq})$.	Ionic hydroxide compounds will dissociate to produce $\text{OH}^-(\text{aq})$.
Ionic compounds are listed as dissociated ions.	Ionic compounds dissociate to some extent in water.
Molecular substances are written in their molecular form; this includes weak acids and weak bases.	The extent of reaction with water is uncertain; therefore, substances are listed in a unionized form.
Metallic ions may be omitted from consideration as possible reactants.	Metallic ions do not have a proton to donate and they do not demonstrate the ability to accept a proton during a chemical reaction.



Self-Check

SC 1. Complete the following chart, which lists the equilibrium position for reactions occurring between substances shown on the “Relative Strengths of Acids and Bases at 298.15 K” table.

Reactants/Position of Reactants	Description of Equilibrium Position	Symbol Used
Hydronium and Hydroxide		
Acid Above Base		
Base Above Acid		

Check your work with the answer in the appendix.



Read

To practise the method of predicting acid-base reactions, work through “Sample problem 16.1” and “Communication example” on pages 729 and 730 of the textbook. You may wish to listen to the “Sample Problem 16.1 Audio Recording.”



For more practice, go to the course multimedia DVD and work through “Predicting Acid-Base Reactions.”



Self-Check

SC 2. Complete “Practice” questions 9–15 on page 731 of the textbook.

Check your work with the answer in the appendix.



Module 8: Lesson 4 Lab—Testing Brønsted-Lowry Reaction Predictions

In this lesson you have learned a method to predict acid-base reactions and their equilibrium position. It is important that all theoretical methods like this one be tested.



Go to the Assignment Booklet, and complete the Pre-Lab section of the “Module 8: Lesson 4 Assignment.” As you view the virtual investigation, record your observations in the appropriate section of your Lesson 4 Assignment.



Go to the course multimedia DVD, and view “Testing Brønsted-Lowry Reaction Predictions.”



Module 8: Lesson 4 Assignment

Go to the Assignment Booklet, and complete the Analysis section of the “Module 8: Lesson 4 Assignment.”

Save your work in your course folder.



Reflect and Connect

At the beginning of this lesson you were asked whether the proposed reaction between nitric acid in acid deposition would be neutralized by carbonate ions present in soil.

Use your understanding of predicting acid-base reactions and equilibrium position to determine whether nitric acid and the other acids responsible for acid deposition (nitrous, sulfuric, and sulfurous acids) would also be neutralized by carbonate ions present in soil.

Save your work in your course folder.

In Lesson 1 you started summaries of your previous learning about acids and bases. Update your summaries to include concepts you learned in this lesson. You may wish to start a new summary for the new skill you learned in this lesson—predicting acid-base reactions. Save a copy of your updated summaries in your course folder. You may be asked to submit a copy to your teacher for feedback.



Lesson Summary

In Lesson 4 you considered the following question:

- How is knowledge of the equilibrium of aqueous systems involved in predicting the outcome of acid-base reactions?

The focus of this lesson has been on developing and testing a method to predict acid-base reactions and the position of their equilibrium. You learned the principles upon which this method of prediction is based, and you tested predictions by completing a virtual investigation. You also considered how an understanding of this method could be applied to investigating the action of soil components against acid deposition.

Knowledge of this method is required to complete the rest of Module 8. In the next few lessons you will further your understanding of acid-base equilibrium by investigating the chemical reaction occurring in a titration and the composition and action of chemical buffers.

Module 8—Acid-Base Equilibrium

Lesson 5—Equilibrium Law and the Strength of Acids and Bases



Get Focused

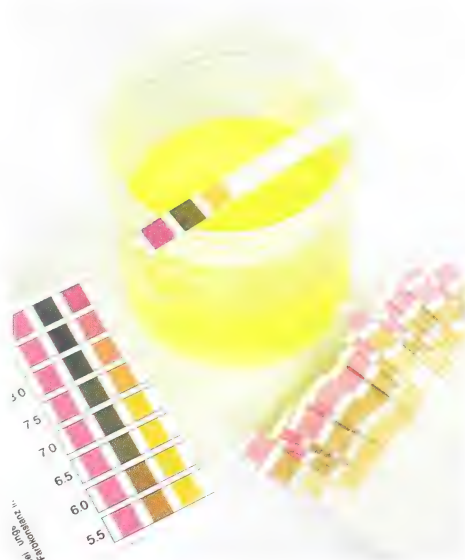
The pH of solutions provides important information about the equilibrium of acidic and basic solutes. In the previous lessons you used solution pH and molar concentration of the solute to calculate a percent ionization. In Module 7 you learned about equilibrium law and K_c , the equilibrium constant. Can K_c be calculated for acidic and basic solutions? What would a difference in the value for the equilibrium constant for two acids or bases tell you about their equilibrium?

You may recall reading in the textbook that the pH of rainwater is normally 5.5. The presence of dissolved carbon dioxide from the atmosphere in the rain droplets results in the formation of carbonic acid, which in turn ionizes to produce hydronium ions. Why is the pH of rainwater expected to be 5.5?

In this lesson you will investigate the use of quantitative techniques to describe the equilibrium of aqueous acids and bases. You will also investigate how the equilibrium influences the pH of both strong and weak acids and bases.

Consider the following questions as you complete Lesson 5:

- What are K_a , K_b , and K_w ?
- How do K_a and K_b explain the position of the equilibrium of aqueous acids and bases?
- How are values for K_a and K_b used to calculate the pH of solutions containing weak acids and bases?



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Module 8: Lesson 5 Assignment

There is no assignment for this lesson. Because of the nature of this lesson, the questions that are not marked by the teacher are very important. You should make an effort to ensure that you are confident with the calculations necessary to complete the work in this lesson before you proceed to Lesson 6.

You must decide what to do with the questions that are not marked by the teacher.

Remember that these questions provide you with the practice and feedback that you need to successfully complete this course. You should respond to all the questions and place those answers in your course folder.



Explore

Your work in Lessons 3 and 4 has helped you add to your knowledge of acids and bases. Because of your work to understand the Brønsted-Lowry theory, you probably find it easier to define acids and bases by what they do in a reaction rather than by their empirical properties.

It is often easy to confuse the term *strength* with *concentration* when talking about acids and bases. *Strength* refers to the extent to which a substance ionizes. For instance, strong acids react 100% to produce hydronium ions. *Concentration* refers to the chemical amount of solute per unit volume of solution. It is possible to have a highly concentrated solution in which a weak acid is the solute.

In Module 7 your exploration of chemical equilibrium involved learning about percent reaction and the equilibrium constant for a system. As you have seen through your studies so far, both values are commonly used to describe the equilibrium position. You may wish to test your memory and write the formulas to calculate percent reaction and the equilibrium constant, and then adapt your calculations for an aqueous acid. How does the similar position of the term $[\text{H}_3\text{O}^+(\text{aq})]$ in both relationships account for the ability to use values for percent reaction and the equilibrium constant to describe equilibrium position?



Read

What would the equilibrium law expression and the equilibrium constant for acidic and basic systems look like? Read pages 737–738 in the textbook to find out. In your readings note that the equilibrium law expression for an acid is slightly different than what you may recall from Module 7. What assumption is being made in order to make this modification?

You may find the “Learning Tip” on page 738 useful in helping you distinguish between the function of a value for percent ionization and a value for K_a , the ionization constant for an acid.



Try This

Work through the “Sample problems” and “Communication examples” on pages 738–741 of the textbook to learn how to calculate K_a .

TR 1. Can you explain why an ICE table was used to solve these problems?

**Read**

In Module 7 you learned that ICE tables are an effective way to represent the changes that occur within a chemical system as the system moves from its initial set-up to equilibrium. You will recall that you only use equilibrium concentrations to calculate the value of the equilibrium constant, whether it be a K_c or a K_a . Completing an ICE table is a good way to sort out the changes that are occurring in a system moving to an equilibrium since, in the action of the forward reaction, reactants will be consumed to make products.

The rows in the ICE table represent the change in the concentration and the equilibrium concentration of each species. The information in these rows are reminders of the chemical processes that occur within the system when components mix and react as the system establishes an equilibrium.

You may notice that, in some cases (e.g., “Sample problem 16.2” and “Communication example 1” on pages 738–739), the concentration of hydronium ions produced as the acid ionizes is quite small. If you look at the first column of the ICE tables in these two textbook examples, you will notice that the “change in concentration” for the acid is so small that the value for the final and initial concentrations are the same. This inaccuracy affects the value for K_a and is the reason that values for K_a , like those that appear in the Chemistry Data Booklet, are often only shown with two significant digits.

As you see in “Communication example 2” on page 739, you cannot always assume that the change in initial concentration of the acid will be small. Therefore, it is a good practice to construct an ICE table when solving these types of problems. Doing so will ensure your calculations have the required accuracy and will demonstrate your complete understanding of the changes that occur as a system develops an equilibrium.

**Self-Check**

SC 1. A 0.011 mol/L solution of oxalic acid has a pH of 2.11. Calculate the K_a for oxalic acid using this information.

Check your work with the answer in the appendix.

**Read****Approximations: Using K_a to Calculate the $[\text{H}_3\text{O}^+(\text{aq})]$ for an Acid**

Because of their incomplete ionization, weak acids can make the calculation to estimate $[\text{H}_3\text{O}^+(\text{aq})]$ or a solution's pH more complex than you might expect. A simple check is often done to determine whether an approximation can be used. As you will see described below, an approximation will be used if the extent of ionization for an acid is low relative to its molar concentration.

If this is the case, then the approximation can be used to reduce the complexity of the calculation required. It is important for you to understand that the reason for running the check to determine whether you can use an approximation is to predict the equilibrium concentration for the acid and to determine whether or not the predicted change in the value of the concentration is significant.

The last series of problems you looked at considered relative amounts of material involved in reactions and whether such values were significant. In “Sample problem 16.2” for example, consider the acetic acid used:

$$\frac{[\text{CH}_3\text{COOH}_{(\text{aq})}]}{K_a} = \frac{1.00 \text{ mol/L}}{1.8 \times 10^{-5}} \\ = 55\,555$$

This value is higher than 1000, so the approximation is valid. As you saw in the example, subtracting the change in concentration was below the precision of the value for the initial amount concentration and was not significant.

Consider a second example, oxalic acid used in SC 1:

$$\frac{[\text{HOOC}\text{COOH}_{(\text{aq})}]}{K_a} = \frac{0.011 \text{ mol/L}}{5.6 \times 10^{-2}} \\ = 0.20$$

This value is significantly below 1000, so the approximation is not valid. This is consistent with the significant change in the concentration of oxalic acid due to its ionization in aqueous solution.

Are you concerned that applying an approximation will decrease the accuracy of your answer? Since the values for K_a shown on the table are only provided to two significant digits, small changes in concentration will not be accounted for in your calculations.

Remember that you must show that you understand when an approximation can be used. It is incorrect to make an assumption about the importance of the change in equilibrium concentration of the acid without showing work to back up your reasoning. The examples you have been asked to look through in the textbook demonstrate how to communicate that you are testing the assumption and how you intend to use that information to solve the problem.



Self-Check

SC 2. Read the information in “Sample problem 16.3” on page 740 in the textbook. Calculate the factor difference between the concentration of methanoic acid and its K_a . Is the approximation valid for methanoic acid, or is the change in concentration during its ionization significant?

Check your work with the answer in the appendix.



Read

Work through “Sample problem 16.3” and “Communication example 3” on pages 740–742 of the textbook.



Self-Check

SC 3. Predict the $[\text{H}_3\text{O}^+(\text{aq})]$ and the pH for a 0.125-mol/L solution of hydrosulfuric acid.

SC 4. Complete “Practice” questions 5 and 6 on page 743 of the textbook.

Check your work with the answer in the appendix.



Try This

Acid-Base Reactions

It is fun to try out the concepts you are learning, and new concepts always seem to make more sense when you try them hands-on. In this simulation you will learn about percent reaction as you try reacting different acids and bases.



Go to the website for your textbook. You will need your username and password, available from your teacher, to access the website. Navigate to Unit 8, Chapter 16, Section 16.3, Extensions. Once there, view the simulation “Acid-Base Reactions.” As you work with the simulation, look for patterns as you choose various reactants. Consider the following questions:

TR 2. Does the K_a help you predict the extent of the reaction between the acid and a base?

TR 3. What does the position on the table indicate about the extent of the reaction?

TR 4. Explain how you could use the simulation to collect information to answer these questions.

Save your answers in your course folder.

**Read****Equilibrium Constant for Bases, K_b**

The analysis of the equilibrium of bases has many similarities to the analyses you have completed so far in this lesson. Read from “Calculating K_b from Amount Concentrations” on page 745 in the textbook to the start of “Calculating $[\text{OH}^-(\text{aq})]$ from K_b ” on page 746.

**Self-Check**

SC 5. Complete “Practice” questions 10–13 on page 746 in the textbook.

Check your work with the answer in the appendix.

**Read**

A relationship exists between the three equilibrium constants you used in this module. Read pages 746–750 in the textbook, and work through the “Sample problems” and “Communication examples” in this section to learn more about the relationship between K_a , K_b , and K_w .

**Self-Check**

SC 6. Complete “Section 16.3” questions 1 and 2 on page 750 in the textbook.

Check your work with the answer in the appendix.

**Reflect and Connect**

In Lesson 1 you started summaries of your previous learning about acids and bases. In this lesson you investigated strong and weak acids and bases using the equilibrium law, and you learned some techniques required to complete pH and pOH calculations for weak acids and bases. You may wish to update your summaries to include concepts you have learned in this lesson.

Save a copy of your updated summaries in your course folder. You may be asked to submit a copy to your teacher.



Lesson Summary

In Lesson 5 you considered the following questions:

- What are K_a , K_b , and K_w ?
- How do K_a and K_b explain the position of the equilibrium of aqueous acids and bases?
- How are values for K_a and K_b used to calculate the pH of solutions containing weak acids and bases?

In this lesson you used your knowledge of equilibrium systems and the ionization constants for acids (K_a), bases (K_b), and water (K_w) to analyze aqueous systems containing either acids or bases. You also used ICE tables and, where applicable, approximations to calculate equilibrium concentrations for weak acids and bases. These calculations enabled you to predict the pH of solutions containing weak acids and bases.

In the next lesson you will continue to investigate the equilibrium relationships of weak acids and bases as you investigate pH curves.

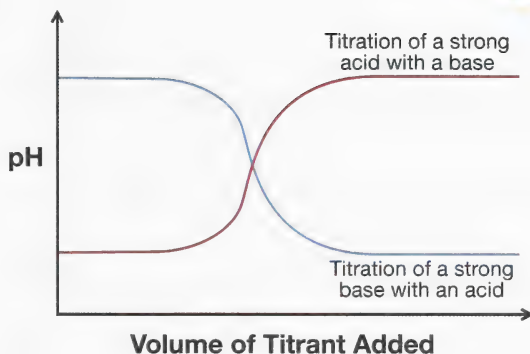
Module 8—Acid-Base Equilibrium

Lesson 6—pH Curves



Get Focused

In Chemistry 20 you completed a titration experiment that involved the stepwise addition of one reactant to another until an endpoint was reached. Another way to set up this experiment would have been to record the change in pH as titrant was added to the test solution. The graph of the data is a pH curve.



In this lesson you will revisit this graph and learn more about its unique shape, including how the shape is influenced by equilibrium. You will also learn about acidic and basic substances that can react multiple times and how their equilibrium influences the shape of a pH curve.

Consider the following questions as you complete Lesson 6:

- What information about acids and bases and their equilibrium is contained on a titration curve?
- What events occur in the reaction of polyprotic acids and bases?



Module 8: Lesson 6 Assignment

As part of your lesson assignment, you will sketch a pH curve. You will be directed to go to the Assignment Booklet to complete the Module 8: Lesson 6 Assignment later in this lesson.

You must decide what to do with the questions that are not marked by the teacher.

Remember that these questions provide you with the practice and feedback that you need to successfully complete this course. You should respond to all the questions and place those answers in your course folder.



Explore

One of the summaries you started back in Lesson 1 focused on titration and titration curves. Can you list the equipment required to perform a titration? What is an indicator used for?

Retrieve the summary you prepared earlier and consider how your knowledge of chemical equilibrium and the equilibrium of acids and bases fits into what you know about titration curves and the process of titration.



Read

Read pages 751–752 in the textbook to learn how equilibrium concepts contribute to the unique shape of a titration curve. You may wish to add this new information on titration and titration curves to the summaries you initiated in Lesson 1.

Read pages 753–754 in the textbook to identify how indicators and the colour changes you used to detect the endpoint of a titration are part of an equilibrium system. As you read this section, consider how the dramatic change in the pH that occurs near the equivalence point of a titration becomes the stress that influences a shift in the equilibrium of the indicator system. This stress eventually leads to the predominance of the other form of the indicator molecule and its characteristic colour.



Self-Check

SC 1. Complete “Practice” questions 1 and 2 on pages 754–755 of the textbook.

Check your work with the answer in the appendix.

**Read**

To this point you have investigated titrations in which one proton transfer occurs. In an earlier lesson in this module you learned that polyprotic acids and bases exist. What would the titration curve for a polyprotic acid or base look like?

Read pages 755–759 in the textbook to learn how the shape of a titration curve for a polyprotic species is influenced by the occurrence of another acid-base reaction. As you read this section, recall the method you used to predict acid-base reactions. At this time also read pages 760–762 to learn how the shape of a pH curve is influenced by the strength of the acid or base being titrated.

**Try This****Sketching Titration Curves**

The distinctive features of a pH curve are as follows:

- starting pH
- appearance of the initial buffering region
- pH at equivalence point
- height of region between buffering regions
- number of equivalence points
- ending pH

Examine the pH curves included on pages 752–757. Describe how the distinctive aspects of titration curves identify the type of substance being titrated.

Save your response in your course folder and submit a copy to your teacher for feedback.

**Self-Check**

SC 2. Complete “Practice” questions 8 and 9 on page 759 of the textbook.

Check your work with the answer in the appendix.

**Module 8: Lesson 6 Assignment**

Go to the Assignment Booklet, and complete the “Module 8: Lesson 6 Assignment.”



Reflect and Connect

Retrieve the summary about titration and titration curves you started in Lesson 1 and referred to earlier in this lesson. You may wish to use this opportunity to add what you have learned in this lesson to your summary.



Lesson Summary

In Lesson 6 you considered the following questions:

- What information about acids and bases and their equilibrium is contained on a titration curve?
- What events occur in the reaction of polyprotic acids and bases?

You expanded upon your understanding of pH curves you began in your previous chemistry course when you graphed the change in pH against the volume of titrant added. Your work here has helped you to read important information on a pH curve, including the following:

- whether the substance being titrated is an acid or a base
- whether the substance being titrated is a strong or weak acid or base
- whether the substance being titrated is monoprotic or polyprotic

With this type of information you are able to comment on the characteristics of the acid or base being studied. If necessary, you can also design appropriate experiments for further study of the acid or base under consideration.

Module 8—Acid-Base Equilibrium

Lesson 7—Buffers



Get Focused

In previous lessons you've been learning about acid deposition and the effect it can have. Why doesn't the soil in Alberta appear to be affected? Is there something about the chemical composition of soil in Alberta that prevents damage from acid deposition?

While it may appear that acid is not having a negative effect, accumulation of acids can adversely affect other systems. Consider the example of tennis. Tennis is all about short bursts of energy. Tennis can be extremely tiring because there is little time to recover between efforts. When you move to return the ball, your body metabolizes energy to allow the muscles to contract. Sometimes the exercise is so intense that your muscles accumulate lactic acid, a by-product of metabolism.

You may recall from previous science courses that the human body is designed to function within a narrow pH range. How does the human body deal with lactic and other acids that may be present? In this lesson you will investigate the properties of buffers as equilibrium systems.



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Consider the following questions as you complete Lesson 7:

- What is a buffer and how does a buffer exhibit the characteristics of an equilibrium system?
- What is buffering capacity?



Module 8: Lesson 7 Assignment

You will be directed to go to the Assignment Booklet to complete the Module 8: Lesson 7 Assignment later in this lesson.

You must decide what to do with the questions that are not marked by the teacher.

Acid-Base Equilibrium

Remember that these questions provide you with the practice and feedback that you need to successfully complete this course. You should respond to all of the questions and place those answers in your course folder.



Explore



Read

In Lesson 6 you labelled certain regions of a pH curve “buffering regions.” What happens in a buffering region compared to other regions in a pH curve? The flattened region of a pH curve indicates that little change in pH occurs despite the addition of titrant. Why so little change?

Read pages 763–765 in the textbook to learn what a buffer is and how this differs from the terms *buffering* and *buffer capacity*, which are also introduced in this section. You may wish to make a chart explaining the similarities and differences between the terms *buffer*, *buffering*, and *buffer capacity* and to save your chart in your course folder for future reference.



Try This

How can your knowledge of equilibrium be used to understand the composition of a chemical buffer?



Go to the website for your textbook. You will need your username and password, available from your teacher, to access the website. Navigate to Unit 8, Chapter 16, Section 16.4, Web Activities. Open the simulation “Preparation of Buffer Solutions.”

Complete the questions in the simulation to identify what kinds of chemical components can be used to form a buffer.



Self-Check

SC 1. Complete “Practice” question 18 on page 766 in the textbook.

Check your work with the answer in the appendix.



Self-Check

How can your knowledge of the equilibrium constant be used to estimate the pH of a buffer solution? Considering that the chemical components in a buffer include a weak acid and its conjugate base and water, would it be possible to determine the pH of a buffer?

SC 2. Use the following information to calculate the pH of a buffer composed of the following two solutions: 0.200 mol/L ethanoic acid and 0.100 mol/L sodium ethanoate.

Check your work with the answer in the appendix.



Module 8: Lesson 7 Assignment

Go to the Assignment Booklet, and complete the “Module 8: Lesson 7 Assignment.”



Reflect on the Big Picture

Acid deposition is a constant stress on soil. Sometimes the quantity of acids deposited overwhelms the buffering capacity of soil, and the soil pH rapidly drops.

Use the Internet to research the effect of loss of buffering capacity in soil on agricultural productivity. What strategies are used to restore the buffering capacity of soil? Do these strategies make sense in light of what you have learned in Module 8?

Save a copy of your research and your response in your course folder.



Lesson Summary

In Lesson 7 you considered the following questions:

- What is a buffer and how does a buffer exhibit the characteristics of an equilibrium system?
- What is buffering capacity?

You learned that buffers are chemical systems composed of a weak acid and its conjugate base. If these components are mixed in distinct proportions, a buffer with a specified pH can be produced.

Buffers act to maintain the pH of a system despite the addition of acid or base but, because they are composed of chemical substances with definite chemical amounts, there is a limit to the ability of a buffer to relieve a stress if excessive chemical amounts are added.

Proceed to the Module 8 Assessment, which also serves as your Unit D Assessment.

Module 8—Acid-Base Equilibrium



Module Summary

In Module 8 you considered the following question:

- How does the equilibrium of acids and bases affect environmental and biological systems?

You investigated the equilibrium of acids and bases. You learned to consider solutions containing acidic and basic solutes as chemical systems in which principles of equilibrium are involved. These principles include extent of reaction, equilibrium position, calculation of equilibrium constant, and determination of solution pH.

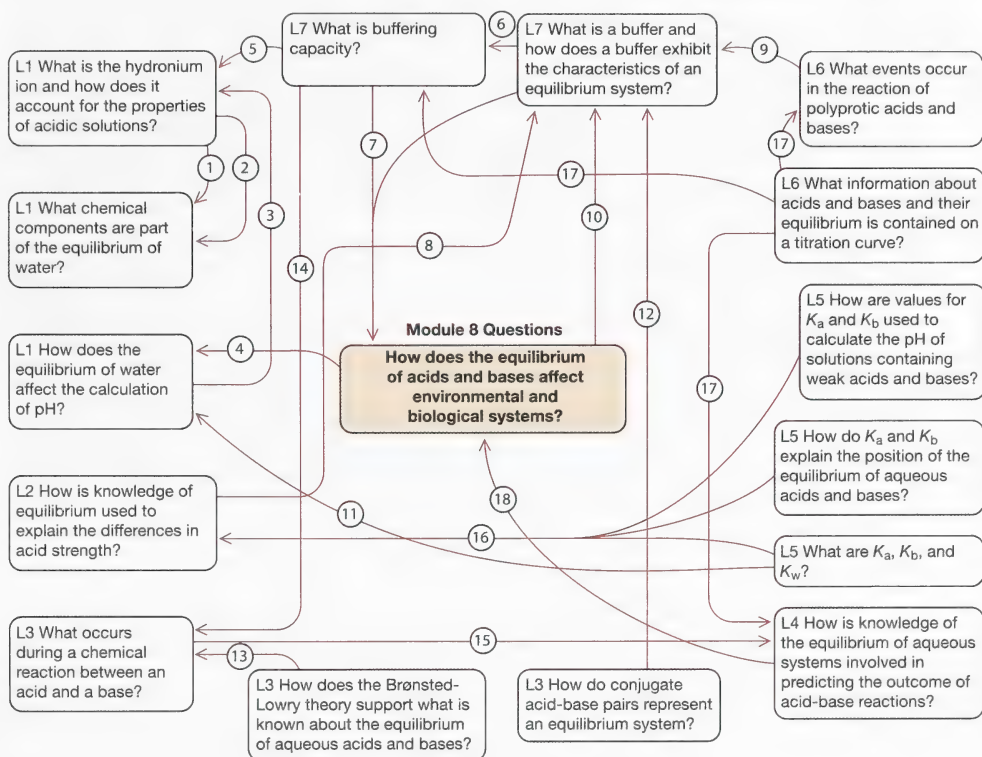
You also continued to apply your understanding of techniques to analyze equilibrium systems, including ICE tables. Throughout this module you considered environmental and biological systems and how the equilibrium of acids and bases is a part of each system's function. This module has also provided you an opportunity to review your understanding of acids and bases developed in your previous chemistry courses and to bring together your understanding of equilibrium to broaden your base of knowledge in this area.

You also learned about buffers, and you learned how these systems demonstrate the principles of equilibrium you studied in Module 7. You examined the importance of buffers in soil and in living systems.

Concept Map or Graphic Organizer

As you worked through Module 8, you may have added information to a concept map or graphic organizer based on the module and lesson questions listed in the Module 8 Concept Organizer. Now is a good time to review the relationships in your concept map or graphic organizer and to try to answer the module and lesson questions.

A sample Module 8 concept map shows one set of possible links between the questions. Remember that this is one possible description only—there are many other correct possibilities. However, if your completed concept map or graphic organizer differs significantly from the sample, you may wish to contact your teacher or to compare your map or organizer with those of other students in your class. This will ensure that your interpretations of lesson materials and your descriptions are accurate.





Module Assessment

This Assessment will serve as your Module 8 Assessment and your Unit D Assessment.

In your study in this module you have used the Relative Strengths of Acids and Bases table. Copies of this table are located on pages 8 and 9 of the Chemistry Data Booklet and on page 829 of the textbook. How is this table designed? In this Assessment you will investigate the structure and listing of information on this table. This Assessment consists of three questions.

1. List, in a general way, the information that is provided on the Relative Strengths of Acids and Bases table.
2. Explain how the information on the Relative Strengths of Acids and Bases table was used to support your work in Unit D. Use examples to support your response.
3. Describe a series of experiments that can be used to confirm the structure and organization of the Relative Strengths of Acids and Bases table. Make sure you include the following information in your response:
 - a description of experiments you would undertake
 - a list of the substances to be tested
 - a description of the tests to be performed and the equipment required to complete these tests
 - a statement of the expected results from the experiments and tests described
 - an explanation of how the expected results would confirm the organization of the Relative Strengths of Acids and Bases table

Scoring Guide

Score	Scoring Description
4 Standard of Excellence	The response is well organized and addresses all the major points of the question using appropriate and clear communication strategies. An appropriate and practical series of experiments is described. Thorough and correct justifications, interpretations, and explanations are provided.
3	The response is organized and addresses most of the major points of the question using appropriate and well formed communication strategies. Appropriate and workable experiments are described. Justifications, interpretations, and explanations are correct .
2 Acceptable Standard	The response is generally organized and addresses most of the major points of the question using logical and reasonable communication strategies. Workable experiments are generally described . Justifications, interpretations, and explanations are correct .
1	The response is poorly organized and addresses few of the major points of the question. Poorly formed arguments and descriptions are used. A plausible experiment is suggested . Justifications, interpretations, and explanations are incorrect or absent from the response.
0	The response is not suitable for a 30-level course.

Unit D Conclusion

In Unit D you studied chemical equilibrium. Module 7 introduced the principles of chemical equilibrium and Module 8 focused on the equilibrium of acids and bases.

Your study in this unit required you to rethink the way you understood chemical systems. By studying chemical equilibrium you have begun to consider what reactions have occurred to get the chemical system to its current equilibrium and how this equilibrium could be altered if conditions were to change.

Throughout Unit D you completed summaries intended to clarify concepts of chemical equilibrium. Your summaries also helped you see how knowledge of equilibrium adds to your understanding of solutions of acids and bases as aqueous chemical systems. Check to make sure you updated your summaries and saved them in your course folder.

If you continue to study chemistry at the post-secondary level, you will be introduced to many areas of chemistry in which considering equilibrium is a regular part of chemical analysis.

You investigated the following module questions:

Module 7

- What is happening in a system at equilibrium?
- How do scientists predict shifts in the equilibrium of a system?

Module 8

- How do Brønsted-Lowry acids and bases illustrate equilibrium?

Because equilibrium allows you to consider the microscopic and macroscopic changes that may occur, you now know that equilibrium is an important tool for studying any chemical system. The chemical industry applies knowledge of chemical equilibrium to manipulate conditions that allow processes to occur. Finally, having qualitatively and quantitatively considered equilibrium, you better understand how strong acids and bases behave differently from weak acids and bases.

You should now be able to

- explain that there is a balance of opposing reactions in chemical equilibrium systems
- determine quantitative relationships in simple equilibrium systems



Self-Check

Go to the course multimedia DVD, and complete the Unit D Diagnostic Self-Check to review the concepts, skills, and knowledge introduced and to assess your understanding of them.

If you have trouble answering any of the questions, try one or more of the following options:

- Go to the lesson(s) indicated for each question. Once there, review the information provided and the relevant sections of the textbook.
- View the hint associated with each question.
- Look at the answer and then work backward to understand the problem and its solution.
- Ask your teacher for help.

You understand how important it is to review material in preparation for any tests your teacher may give you and in solidifying your understanding of new concepts. Completing these questions is just one part of your review. Program of Studies information is associated with each question. You may wish to use this information to identify areas in which you need to pay particular attention. You may also answer the questions that appear at the end of each chapter and unit in the textbook.

You have already reviewed your Modules 7 and 8 concept maps or graphic organizers. Now is a good time to give them another look and add any new connections or ideas to them.

Unit D Assessment



Unit Assessment

Your Unit D Assessment was included in your Module 8 Assessment.

Chemistry 30 Conclusion**Course Conclusion**

Congratulations, you have completed Chemistry 30 Learn EveryWare!



In preparation for your Chemistry 30 Diploma Examination you may wish to revisit the Diagnostic Self-Check activities that appear at the end of each unit on the course multimedia DVD.

- Unit A Diagnostic Self-Check
- Unit B Diagnostic Self-Check
- Unit C Diagnostic Self-Check
- Unit D Diagnostic Self-Check

As you work through these Self-Check questions you may wish to spend a little more time on the extra information provided. For example, read the Program of Studies objectives, and then think about the other questions you have seen in this course that assess your knowledge of the same Program of Studies objectives. What kinds of questions might appear on the Diploma Examination to test your understanding of each objective?

Another strategy to prepare for your diploma examination may be to prepare a “can-do” list for each of the objectives in the Chemistry 30 Program of Studies. Simply go through the list of objectives and determine whether you “can do” each of the knowledge and skills objectives listed. For each of the objectives, you may wish to include an example of a question, exercise, or lab in which you demonstrated your understanding of the objective.

Contact your teacher to see whether there will be a final test or assessment for you to complete prior to your Diploma Examination. Good luck!

Appendix



Module 8 Self-Check Answers

Contact your teacher if your answers vary significantly from the answers provided here.

Lesson 1

SC 1.

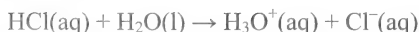
Practice 1.

$$\begin{aligned}
 [\text{OH}^- (\text{aq})] &= \frac{K_w}{[\text{H}_3\text{O}^+ (\text{aq})]} \\
 &= \frac{1.0 \times 10^{-14}}{4.40 \times 10^{-3} \text{ mol/L}} \\
 &= 2.3 \times 10^{-12} \text{ mol/L}
 \end{aligned}$$

Practice 2.

$$\begin{aligned}
 [\text{H}_3\text{O}^+ (\text{aq})] &= \frac{K_w}{[\text{OH}^- (\text{aq})]} \\
 &= \frac{1.0 \times 10^{-14}}{0.299 \text{ mmol/L}} \\
 &= 3.3 \times 10^{-11} \text{ mol/L}
 \end{aligned}$$

Practice 3.



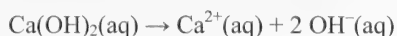
Since $\text{HCl}(\text{aq})$ is a strong acid, it reacts 100% with water; therefore,

$$[\text{H}_3\text{O}^+(\text{aq})] = [\text{HCl}(\text{aq})]$$

$$\begin{aligned}
 [\text{HCl}(\text{aq})] &= \frac{\left(\frac{0.37 \text{ g}}{(1.01 \text{ g/mol} + 35.45 \text{ g/mol})} \right)}{\left(250 \text{ mL} \times \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) \right)} \\
 &= 0.041 \text{ mol/L}
 \end{aligned}$$

$$[\text{H}_3\text{O}^+(\text{aq})] = 0.041 \text{ mol/L}$$

$$\begin{aligned}
 [\text{OH}^-(\text{aq})] &= \frac{K_w}{[\text{H}_3\text{O}^+(\text{aq})]} \\
 &= \frac{1.0 \times 10^{-14}}{0.041 \text{ mol/L}} \\
 &= 2.5 \times 10^{-13} \text{ mol/L}
 \end{aligned}$$

Practice 4.

$$\begin{aligned}
 [\text{OH}^-(\text{aq})] &= \frac{6.9 \text{ mmol/L} \times 2_{\text{OH}^-}}{1_{\text{Ca}(\text{OH})_2}} \\
 &= 14 \text{ mmol/L}
 \end{aligned}$$

$$\begin{aligned}
 [\text{H}_3\text{O}^+(\text{aq})] &= \frac{K_w}{[\text{OH}^-(\text{aq})]} \\
 &= \frac{1.0 \times 10^{-14}}{14 \text{ mmol/L}} \\
 &= 7.2 \times 10^{-13} \text{ mol/L}
 \end{aligned}$$

SC 2.**Practice 7.**

a.

	$[\text{H}_3\text{O}^+(\text{aq})]$ (mol/L)	$[\text{OH}^-(\text{aq})]$ (mol/ L)	pH	pOH
Oranges	5.5×10^{-3}	1.8×10^{-12} $\left(= \frac{1.0 \times 10^{-14}}{5.5 \times 10^{-3}} \right)$	2.3 $(= -\log(5.5 \times 10^{-3}))$	11.74 $(= 14.00 - 2.26)$
Asparagus	3×10^{-9} $\left(= \frac{1.0 \times 10^{-14}}{3 \times 10^{-6}} \right)$	3×10^{-6} $(= 10^{-5.6})$	8.4 $(= 14.00 - 5.6)$	5.6

Acid-Base Equilibrium

Olives	5×10^{-4} $\left(= \frac{1.0 \times 10^{-14}}{2.0 \times 10^{-11}} \right)$	2.0×10^{-11}	3.30	10.70 (= 14.00 – 3.30)
Blackberries	4×10^{-4} (= $10^{-3.4}$)	3×10^{-11} (= $10^{-10.6}$)	3.4 (= 14.00 – 10.6)	10.6

b. The oranges would taste most sour because they have the lowest pH.

Practice 8.



Since NaOH(aq) is highly soluble in water, it completely dissociates. Therefore,

$$[\text{OH}^-(\text{aq})] = [\text{NaOH(aq)}]$$

$$[\text{NaOH(aq)}] = \frac{\left(\frac{26 \text{ g}}{(22.99 \text{ g/mol} + 1.01 \text{ g/mol} + 16.00 \text{ g/mol})} \right)}{\left(150 \text{ mL} \times \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) \right)}$$

$$= 4.3 \text{ mol/L}$$

$$[\text{OH}^-(\text{aq})] = 4.3 \text{ mol/L}$$

$$\text{pOH} = -\log(4.3 \text{ mol/L})$$

$$= -0.64$$

$$\text{pH} = 14.00 - (-0.64)$$

$$= 14.64$$

SC 3.

- Both solutions, sodium chloride and hydrochloric acid, contain dissociated chloride ions. Therefore, there will be no reaction when the two solutions are mixed.
- $\text{NaCl(aq)} \rightleftharpoons \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq})$
- $K_c = [\text{Na}^+(\text{aq})][\text{Cl}^-(\text{aq})]$. Since the sodium chloride solution is saturated, all substances in the system have a constant concentration.
- When the concentrated hydrochloric acid was added, a white precipitate formed. It appears that the addition of HCl(aq) shifts the equilibrium to the left, favouring the reverse reaction and allowing the formation of NaCl(s) .

- e. The addition of HCl(aq) adds $\text{Cl}^- (\text{aq})$ and some water to the system but does not add $\text{Na}^+ (\text{aq})$. Therefore, the concentration of sodium ions decreases when HCl(aq) is added. This could be a stress on the system and, as predicted by Le Châtelier's principle, the forward reaction would be favoured (to increase the concentration of $\text{Na}^+ (\text{aq})$).

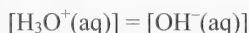
The result of this stress is a further increase in the concentration of chloride ions to a level that exceeds the concentration that is possible to remain dissolved. This situation favours the reverse reaction, and the precipitate of NaCl(s) forms.

- f. The concentration of chloride ions in the concentrated hydrochloric acid must have been higher than the concentration of chloride ions in the saturated sodium chloride solution. Any other possibility would not have resulted in a situation that increased the concentration of $\text{Cl}^- (\text{aq})$ past its maximum solubility.

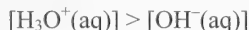
Lesson 2

Section 16.1 1.

- a. If a solution is neutral, then



- b. If a solution is acidic, then



- c. If a solution is basic, then

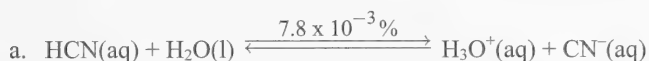


Section 16.1 2.

Conductivity: This test would detect the degree of ionization. Strong acids will ionize to a greater extent than weak acids and, therefore, would produce solutions with a higher conductivity.

pH: Since strong acids ionize to a greater extent than weak acids, solutions of equimolar solute concentrations of strong acids will have lower pHs.

Section 16.1 4.



Acid-Base Equilibrium

b.

$$\begin{aligned}\left[\text{H}_3\text{O}^+(\text{aq})\right] &= 0.10 \text{ mol/L} \left(\frac{7.8 \times 10^{-3}}{100} \right) \\ &= 7.8 \times 10^{-6} \text{ mol/L}\end{aligned}$$

$$\begin{aligned}\text{pH} &= -\log(7.8 \times 10^{-6} \text{ mol/L}) \\ &= 5.11\end{aligned}$$

Section 16.1 6.

$$\begin{aligned}\left[\text{H}_3\text{O}^+(\text{aq})\right] &= \frac{1.0 \times 10^{-14}}{2.5 \times 10^{-7} \text{ mol/L}} \\ &= 4.0 \times 10^{-8} \text{ mol/L}\end{aligned}$$

$$\begin{aligned}\text{pH} &= -\log(4.0 \times 10^{-8} \text{ mol/L}) \\ &= 7.40\end{aligned}$$

Section 16.1 7.

$$\begin{aligned}\left[\text{H}_3\text{O}^+(\text{aq})\right] &= 10^{-5.6} \\ &= 3 \times 10^{-6} \text{ mol/L}\end{aligned}$$

Lesson 3

SC 1.

Practice 1.

- According to the Arrhenius theory, acids are substances that ionize in water to form aqueous hydrogen ions.
- According to the modified Arrhenius theory, acids are substances that react with water to form aqueous hydronium ions.
- According to the Brønsted-Lowry concept, acids are substances that donate a hydrogen ion (proton) during chemical reactions.

Practice 2.

Both definitions predict the hydroxide ions as a product of the processes involved but differ in explaining how the hydroxide ion is produced. The Arrhenius theory explains the ion's production as a result of dissociation of the solute. The Brønsted-Lowry concept requires that a

substance react with water; in the resulting transfer of a proton from the water molecule to the solute, a hydroxide ion is produced.

Practice 3.

- $\text{HF}(\text{aq})$ acid, $\text{SO}_3^{2-}(\text{aq})$ base
- $\text{CO}_3^{2-}(\text{aq})$ base, $\text{CH}_3\text{COOH}(\text{aq})$ acid
- $\text{H}_3\text{PO}_4(\text{aq})$ acid, $\text{OCl}^-(\text{aq})$ base
- $\text{HCO}_3^-(\text{aq})$ base, $\text{HSO}_4^-(\text{aq})$ acid

Practice 4.

- $$\begin{array}{ccc} \text{HSO}_3^-(\text{aq}) + \text{OH}^-(\text{aq}) & \rightleftharpoons & \text{SO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \\ \text{acid} & & \text{base} \end{array}$$
- $$\begin{array}{ccc} \text{HSO}_3^-(\text{aq}) + \text{H}_3\text{O}^+(\text{aq}) & \rightleftharpoons & \text{H}_2\text{SO}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \\ \text{base} & & \text{acid} \end{array}$$

Practice 5.

The Brønsted-Lowry concept allows for the ability to describe the reaction between acids and bases and to explain how some substances could behave as amphiprotic species.

SC 2.

Practice 7.

- $\text{HCO}_3^-(\text{aq})$, $\text{CO}_3^{2-}(\text{aq})$ and $\text{S}^{2-}(\text{aq})$, $\text{HS}^-(\text{aq})$
- $\text{H}_2\text{CO}_3(\text{aq})$, $\text{HCO}_3^-(\text{aq})$ and $\text{OH}^-(\text{aq})$, $\text{H}_2\text{O}(\text{l})$
- $\text{HSO}_4^-(\text{aq})$, $\text{SO}_4^{2-}(\text{aq})$ and $\text{HPO}_4^{2-}(\text{aq})$, $\text{H}_2\text{PO}_4^-(\text{aq})$
- $\text{H}_2\text{O}(\text{l})$, $\text{H}_3\text{O}^+(\text{aq})$ and $\text{H}_2\text{O}(\text{l})$, $\text{OH}^-(\text{aq})$

Practice 8.

$\text{HCO}_3^-(\text{aq})$, $\text{CO}_3^{2-}(\text{aq})$ and $\text{H}_2\text{CO}_3(\text{aq})$, $\text{HCO}_3^-(\text{aq})$

Lesson 4

SC 1.

Reactants/Position of Reactants	Description of Equilibrium Position	Symbol Used
Hydronium and Hydroxide	quantitative	$\rightarrow$
Acid Above Base	products favoured (greater than 50%)	$\xrightleftharpoons{>50\%}$
Base Above Acid	reactants favoured (less than 50%)	$\xrightleftharpoons{<50\%}$

SC 2.

Practice 9.

Step 1: $\text{HF(aq)}, \text{Na}^+(\text{aq}), \text{SO}_4^{2-}, \text{H}_2\text{O(l)}$

Step 2: $\text{HF}^{\text{A}}(\text{aq}), \text{Na}^+(\text{aq}), \text{SO}_4^{\text{B}2-}(\text{aq}), \text{H}_2\text{O}^{\text{A}}(\text{l})$

Step 3: Strongest acid: HF(aq)
Strongest base: $\text{SO}_4^{2-}(\text{aq})$

Step 4: $\text{HF(aq)} + \text{SO}_4^{2-}(\text{aq}) \rightleftharpoons \text{F}^-(\text{aq}) + \text{HSO}_4^-(\text{aq})$

Step 5: $\text{HF(aq)} + \text{SO}_4^{2-}(\text{aq}) \xrightleftharpoons{<50\%} \text{F}^-(\text{aq}) + \text{HSO}_4^-(\text{aq})$

The predominant reaction is equilibrium favouring the reactants.

Practice 10.

Step 1: $\text{H}_3\text{O}^+(\text{aq}), \text{ClO}_4^-(\text{aq}), \text{Na}^+(\text{aq}), \text{OH}^-(\text{aq}), \text{H}_2\text{O(l)}$

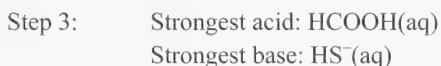
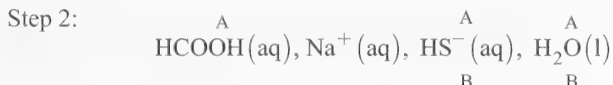
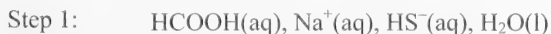
Step 2: $\text{H}_3\text{O}^{\text{A}}(\text{aq}), \text{ClO}_4^{\text{B}-}(\text{aq}), \text{Na}^+(\text{aq}), \text{OH}^{\text{B}-}(\text{aq}), \text{H}_2\text{O}^{\text{A}}(\text{l})$

Step 3: Strongest acid: $\text{H}_3\text{O}^+(\text{aq})$
Strongest base: $\text{OH}^-(\text{aq})$



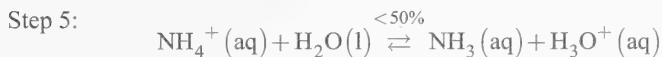
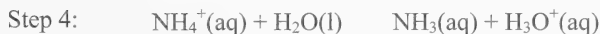
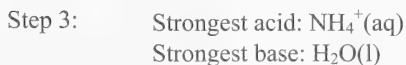
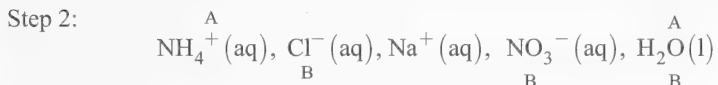
The predominant reaction is a quantitative reaction.

Practice 11.



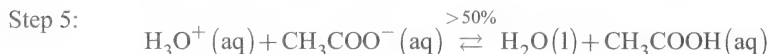
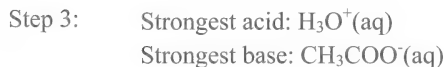
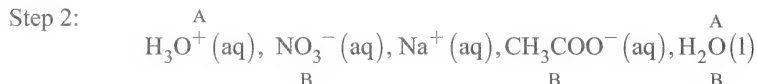
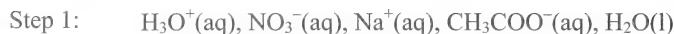
The predominant reaction is equilibrium favouring the products.

Practice 12.



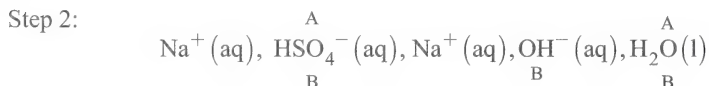
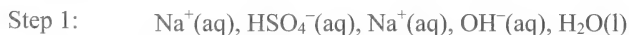
The predominant reaction is equilibrium favouring the reactants.

Practice 13.



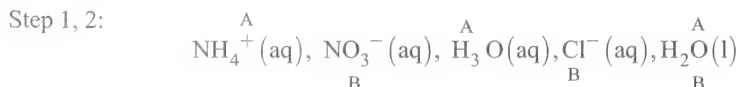
Equilibrium favouring the products is predicted. Empirical testing confirms a quantitative reaction.

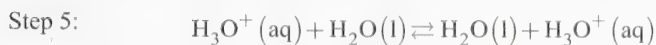
Practice 14.



The predominant reaction is equilibrium favouring the products.

Practice 15.





The predominant reaction is no change; reactants and products are identical.

Lesson 5

SC 1.



$$K_a = \frac{[\text{HOOC}\text{COO}^-(\text{aq})][\text{H}_3\text{O}^+(\text{aq})]}{[\text{HOOC}\text{COOH}(\text{aq})]}$$

At equilibrium,

$$\begin{aligned} [\text{H}_3\text{O}^+(\text{aq})] &= 10^{-2.11} \\ &= 0.0078 \text{ mol/L} \end{aligned}$$

Amount Concentration	$[\text{HOOC}\text{COOH}(\text{aq})]$ (mol/L)	$[\text{HOOC}\text{COO}^-(\text{aq})]$ (mol/L)	$[\text{H}_3\text{O}^+(\text{aq})]$ (mol/L)
Initial	0.011	0	0
Change	-0.0078 $\left(= 0.0078 \times \frac{1_{\text{HOOC}\text{COOH}}}{1_{\text{H}_3\text{O}^+}} \right)$	$+0.0078$ $\left(= 0.0078 \times \frac{1_{\text{HOOC}\text{COOH}^-}}{1_{\text{H}_3\text{O}^+}} \right)$	$+0.0078$
Equilibrium	0.003	0.0078	0.0078

$$\begin{aligned} K_a &= \frac{[0.0078 \text{ mol/L}][0.0078 \text{ mol/L}]}{[0.003 \text{ mol/L}]} \\ &= 1.9 \times 10^{-2} \end{aligned}$$

According to the data, the K_a for oxalic acid is 1.9×10^{-2} .

SC 2.

$$\frac{[\text{HCOOH}_{(\text{aq})}]}{K_{\text{a}}} = \frac{0.200 \text{ mol/L}}{1.8 \times 10^{-4}} \\ = 1111$$

This value is higher than 1000, so the approximation is valid. As a result, the change in concentration is not significant and the initial concentration of methanoic acid, 0.200 mol/L, can be used in further calculations as required.

SC 3.

Test the following assumption:

$$\frac{[\text{H}_2\text{S}(\text{aq})]}{K_{\text{a}}} = \frac{0.125 \text{ mol/L}}{8.9 \times 10^{-8}} \\ = 1\,404\,494$$

This value is greater than 1000, so the approximation is valid. As a result, the change in the concentration of hydrosulfuric acid due to its ionization is very small, and the equilibrium concentration of hydrosulfuric acid is assumed to be 0.125 mol/L.

Let $x = [\text{H}_3\text{O}^+(\text{aq})]$. Therefore,

$$x = [\text{HS}^-(\text{aq})]$$

$$K_{\text{a}} = \frac{[\text{HS}^-(\text{aq})][\text{H}_3\text{O}^+(\text{aq})]}{[\text{H}_2\text{S}(\text{aq})]}$$

$$8.9 \times 10^{-8} = \frac{[x][x]}{[0.125 \text{ mol/L}]}$$

$$x^2 = 8.9 \times 10^{-8} [0.125 \text{ mol/L}]$$

$$x = \sqrt{8.9 \times 10^{-8} [0.125 \text{ mol/L}]}$$

$$x = 1.1 \times 10^{-4} \text{ mol/L}$$

$$\text{pH} = -\log(1.1 \times 10^{-4} \text{ mol/L}) \\ = 3.98$$

SC 4.

Practice 5.

a.

$$\begin{aligned}
 \% \text{ reaction} &= \left(\frac{[\text{H}_3\text{O}^+(\text{aq})]}{\text{Propanoic acid}(\text{aq})} \right) \times 100 \\
 &= \left(\frac{1.16 \times 10^{-3} \text{ mol/L}}{0.100 \text{ mol/L}} \right) \times 100 \\
 &= 1.16\%
 \end{aligned}$$

b.

Amount Concentration	$[\text{C}_2\text{H}_5\text{COOH}(\text{aq})]$ (mol/L)	$[\text{C}_2\text{H}_5\text{COO}^-(\text{aq})]$ (mol/L)	$[\text{H}_3\text{O}^+(\text{aq})]$ (mol/L)
Initial	0.100	0	0
Change	-1.16×10^{-3} $\left(= 1.16 \times 10^{-3} \times \frac{1_{\text{C}_2\text{H}_5\text{COOH}}}{1_{\text{H}_3\text{O}^+}} \right)$	$+1.16 \times 10^{-3}$ $\left(= 1.16 \times 10^{-3} \times \frac{1_{\text{C}_2\text{H}_5\text{COO}^-}}{1_{\text{H}_3\text{O}^+}} \right)$	$+1.16 \times 10^{-3}$
Equilibrium	0.099	1.16×10^{-3}	1.16×10^{-3}

$$\begin{aligned}
 K_a &= \frac{[\text{C}_2\text{H}_5\text{COO}^-(\text{aq})][\text{H}_3\text{O}^+(\text{aq})]}{[\text{C}_2\text{H}_5\text{COOH}(\text{aq})]} \\
 K_a &= \frac{[1.16 \times 10^{-3} \text{ mol/L}][1.16 \times 10^{-3} \text{ mol/L}]}{[0.099 \text{ mol/L}]} \\
 &= 1.36 \times 10^{-5}
 \end{aligned}$$

c. As with all equilibrium constants, the value is only constant at the temperature at which the data were collected.

Practice 6.

a. $[\text{H}_3\text{O}^+(\text{aq})] = 10^{-2.43}$

$$= 3.7 \times 10^{-3} \text{ mol/L}$$

$$\% \text{ reaction} = \left(\frac{[\text{H}_3\text{O}^+(\text{aq})]}{[\text{2-hydroxypropanoic acid}(\text{aq})]} \right) \times 100$$

$$= \left(\frac{3.7 \times 10^{-3} \text{ mol/L}}{0.10 \text{ mol/L}} \right) \times 100$$

$$= 3.7\%$$

b.

Amount Concentration	$[\text{C}_2\text{H}_5\text{OCOOH}(\text{aq})]$ (mol/L)	$[\text{C}_2\text{H}_5\text{OCOO}^-(\text{aq})]$ (mol/L)	$[\text{H}_3\text{O}^+(\text{aq})]$ (mol/L)
Initial	0.10	0	0
Change	-3.7×10^{-3} $\left(= 3.7 \times 10^{-3} \times \frac{1_{\text{C}_2\text{H}_5\text{OCOOH}}}{1_{\text{H}_3\text{O}^+}} \right)$	$+3.7 \times 10^{-3}$ $\left(= 3.7 \times 10^{-3} \times \frac{1_{\text{C}_2\text{H}_5\text{OCOO}^-}}{1_{\text{H}_3\text{O}^+}} \right)$	$+3.7 \times 10^{-3}$
Equilibrium	0.10 $\left[= 0.10 - 3.7 \times 10^{-3} \right]$ $= 0.096$	3.7×10^{-3}	3.7×10^{-3}

$$K_a = \frac{[\text{C}_2\text{H}_5\text{OCOO}^-(\text{aq})][\text{H}_3\text{O}^+(\text{aq})]}{[\text{C}_2\text{H}_5\text{OCOOH}(\text{aq})]}$$

$$K_a = \frac{[3.7 \times 10^{-3} \text{ mol/L}][3.7 \times 10^{-3} \text{ mol/L}]}{[0.10 \text{ mol/L}]}$$

$$= 1.4 \times 10^{-4}$$

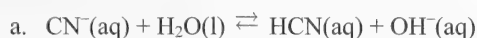
- c. The addition of the hydroxyl group increases the K_a by a factor of 10. Therefore 2-hydroxypropanoic acid ionizes ten times more than propanoic acid.

SC 5.

Practice 10.

Empirical Property	Expected Result with a Strong Base	Expected Result with a Weak Base
conductivity	strong electrolyte	weak electrolyte
pH	high	lower, but above 7
rate of reaction	higher	lower

Practice 11.



$$K_b = \frac{[\text{HCN}(\text{aq})][\text{OH}^-(\text{aq})]}{[\text{CN}^-(\text{aq})]}$$



$$K_b = \frac{[\text{HSO}_4^-(\text{aq})][\text{OH}^-(\text{aq})]}{[\text{SO}_4^{2-}(\text{aq})]}$$

Practice 12.

Amount Concentration	$[\text{C}_2\text{H}_5\text{COO}^-(\text{aq})]$ (mol/L)	$[\text{C}_2\text{H}_5\text{COOH}(\text{aq})]$ (mol/L)	$[\text{OH}^-(\text{aq})]$ (mol/L)
Initial	0.157	0	0
Change	-1.1×10^{-5} $\left(= 1.1 \times 10^{-5} \times \frac{1_{\text{C}_2\text{H}_5\text{COO}^-}}{1_{\text{OH}^-}} \right)$	$+1.1 \times 10^{-5}$ $\left(= 1.1 \times 10^{-5} \times \frac{1_{\text{C}_2\text{H}_5\text{COOH}}}{1_{\text{OH}^-}} \right)$	$+1.1 \times 10^{-5}$
Equilibrium	0.157 $\left(= 0.157 - 1.1 \times 10^{-5} \right)$ $= 0.157\,698\,9$, an insignificant change	1.1×10^{-5}	1.1×10^{-5}

Acid-Base Equilibrium

$$K_b = \frac{[\text{C}_2\text{H}_5\text{COOH}(\text{aq})][\text{OH}^-(\text{aq})]}{[\text{C}_2\text{H}_5\text{COO}^-(\text{aq})]}$$

$$K_b = \frac{[1.1 \times 10^{-5} \text{ mol/L}][1.1 \times 10^{-5} \text{ mol/L}]}{[0.157 \text{ mol/L}]}$$

$$= 7.7 \times 10^{-10}$$

Practice 13.

$$\text{pOH} = 14.00 - 8.81$$

$$= 5.19$$

$$[\text{OH}^-(\text{aq})] = 10^{-5.19}$$

$$= 6.5 \times 10^{-6} \text{ mol/L}$$

Amount Concentration	$[\text{C}_6\text{H}_5\text{NH}_2(\text{aq})]$ (mol/L)	$[\text{C}_6\text{H}_5\text{NH}_3^+(\text{aq})]$ (mol/L)	$[\text{OH}^-(\text{aq})]$ (mol/L)
Initial	0.10	0	0
Change	-6.5×10^{-6} $\left(= 6.5 \times 10^{-6} \times \frac{1_{\text{C}_6\text{H}_5\text{NH}_2}}{1_{\text{OH}^-}} \right)$	$+6.5 \times 10^{-6}$ $\left(= 6.5 \times 10^{-6} \times \frac{1_{\text{C}_6\text{H}_5\text{NH}_3^+}}{1_{\text{OH}^-}} \right)$	$+6.5 \times 10^{-6}$
Equilibrium	0.10 (an insignificant change)	6.5×10^{-6}	6.5×10^{-6}

$$K_b = \frac{[6.5 \times 10^{-6} \text{ mol/L}][6.5 \times 10^{-6} \text{ mol/L}]}{[0.10 \text{ mol/L}]}$$

$$= 4.2 \times 10^{-10}$$

SC 6.

Section 16.3 1.



Approximation

$$\frac{[\text{Cod}(\text{aq})]}{K_b} = \frac{0.020 \text{ mol/L}}{1.73 \times 10^{-6}}$$

$$= 11\,560$$

This is significantly higher than 1000; so the assumption holds, and the change in concentration in codeine is insignificant.

$$K_b = \frac{[\text{Cod}^+(\text{aq})][\text{OH}^-(\text{aq})]}{[\text{Cod}(\text{aq})]}$$

$$\text{Let } x = [\text{OH}^-(\text{aq})]$$

$$= [\text{Cod}^+(\text{aq})]$$

$$1.73 \times 10^{-6} = \frac{[x][x]}{[0.020 \text{ mol/L}]}$$

$$x = \sqrt{1.73 \times 10^{-6} [0.020 \text{ mol/L}]}$$

$$x = 1.9 \times 10^{-4} \text{ mol/L}$$

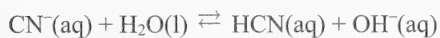
$$[\text{OH}^-(\text{aq})] = 1.9 \times 10^{-4} \text{ mol/L}$$

$$\text{pOH} = -\log(1.9 \times 10^{-4} \text{ mol/L})$$

$$= 3.73$$

$$\text{pH} = 14.00 - 3.73$$

$$= 10.27$$

Section 16.3 2.

$$K_b \text{ of cyanide ion} = \frac{K_w}{K_a \text{ HCN}}$$

$$= \frac{1.00 \times 10^{-14}}{6.2 \times 10^{-10}}$$

$$= 1.6 \times 10^{-5}$$

Acid-Base Equilibrium

Approximation

$$\frac{0.18 \text{ mol/L}}{1.6 \times 10^{-5}} = 11\,160$$

This is greater than 1000, so the change in concentration of CN^- is not significant.

$$K_b = \frac{[\text{HCN}(\text{aq})][\text{OH}^-(\text{aq})]}{[\text{CN}^-(\text{aq})]}$$

$$\begin{aligned}\text{Let } x &= [\text{OH}^-(\text{aq})] \\ &= [\text{HCN}(\text{aq})]\end{aligned}$$

$$\begin{aligned}1.6 \times 10^{-5} &= \frac{[x][x]}{0.18 \text{ mol/L}} \\ x &= \sqrt{1.61 \times 10^{-5} (0.18 \text{ mol/L})} \\ x &= 0.0017 \text{ mol/L}\end{aligned}$$

$$[\text{OH}^-(\text{aq})] = 0.0017 \text{ mol/L}$$

$$\begin{aligned}\text{pOH} &= -\log(0.0017 \text{ mol/L}) \\ &= 2.77\end{aligned}$$

$$\begin{aligned}\text{pH} &= 14.00 - 2.77 \\ &= 11.23\end{aligned}$$

Lesson 6

SC 1.

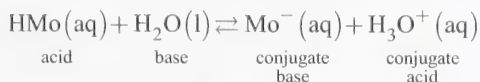
Practice 1.

- The slope of the buffering region is relatively flat (slope = nearly 0).
- The buffering region is where the concentration of the reactant in the test solution is reducing as a result of being converted chemically due to the addition of titrant.
- Near the equivalence point, a rapid change in pH occurs; therefore, the slope of the graph is very large.
- Nearing the equivalence point, the chemical amount of titrant added is almost equal to the chemical amount of reactant in the test solution. The dramatic change in pH witnessed is

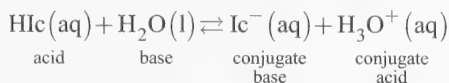
a direct result of the chemical conversion of the reactant in the test solution and its decreased influence on the properties of the solution. Past the equivalence point, titrant is in excess and the pH change is due to the accumulation of titrant and its influence on the properties of the resulting solution.

Practice 2.

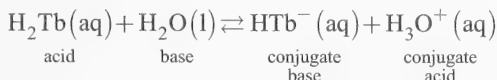
a.



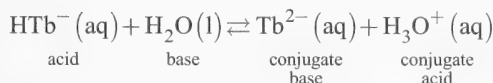
b.



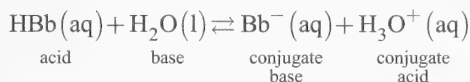
c.



d.



e.



SC 2.

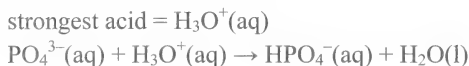
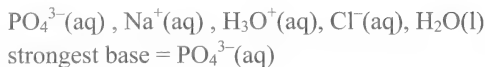
Practice 8.

- around pH 9
- phenolphthalein (pink endpoint), thymol blue (blue endpoint)
- $\text{CH}_3\text{COOH}(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{CH}_3\text{COO}^-(\text{aq}) + \text{H}_2\text{O}(\text{l})$
- The addition of titrant converts a small quantity of acetic acid into acetate ion, which has basic properties. As a result, there is a jump in pH. As titrant continues to be added, there is only a gradual change in the removal of acetic acid particles and appearance of acetate ions; so the pH change is less dramatic.

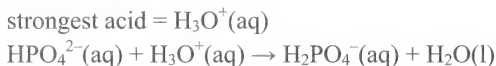
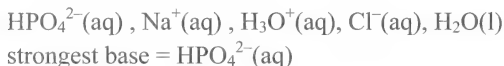
Practice 9.

a. and b.

Two reactions occur:



Since the first reaction was quantitative and titrant continues to be added, the possible reactants include:

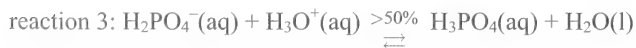
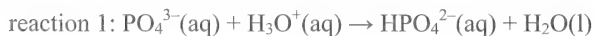


The second reaction is quantitative as well.

The positions of $\text{H}_2\text{PO}_4^-(\text{aq})$ and $\text{H}_3\text{O}^+(\text{aq})$ do not indicate a quantitative reaction would occur in a third reaction, although the reaction would favour the products.

Since there are two quantitative reactions, there would be two equivalence points on a pH curve.

c.



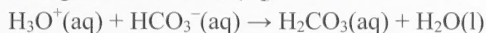
Lesson 7**SC 1.****Practice 18.**

- a. Hydrochloric acid is a strong acid; therefore, it is written as $\text{H}_3\text{O}^+(\text{aq})$ and $\text{Cl}^-(\text{aq})$.

species: $\text{H}_2\text{CO}_3(\text{aq})$, $\text{HCO}_3^-(\text{aq})$, $\text{H}_2\text{O}(\text{l})$, $\text{H}_3\text{O}^+(\text{aq})$, $\text{Cl}^-(\text{aq})$

strongest acid: $\text{H}_3\text{O}^+(\text{aq})$

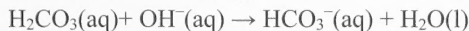
strongest base: $\text{HCO}_3^-(\text{aq})$



- b. species: $\text{H}_2\text{CO}_3(\text{aq})$, $\text{HCO}_3^-(\text{aq})$, $\text{H}_2\text{O}(\text{l})$, $\text{OH}^-(\text{aq})$, $\text{Na}^+(\text{aq})$

strongest acid: $\text{H}_2\text{CO}_3(\text{aq})$

strongest base: $\text{OH}^-(\text{aq})$

**SC 2.**

$$K_a = \frac{[\text{CH}_3\text{COO}^-(\text{aq})][\text{H}_3\text{O}^+(\text{aq})]}{[\text{CH}_3\text{COOH}(\text{aq})]}$$

$$1.8 \times 10^{-5} = \frac{[0.100 \text{ mol/L}][\text{H}_3\text{O}^+(\text{aq})]}{[0.200 \text{ mol/L}]}$$

$$[\text{H}_3\text{O}^+(\text{aq})] = 3.6 \times 10^{-5} \text{ mol/L}$$

$$\begin{aligned} \text{pH} &= -\log(3.6 \times 10^{-5} \text{ mol/L}) \\ &= 4.44 \end{aligned}$$

